DOCUMENT RESUME

ED 166 038 SE 026 370

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TITLE An Introduction to Chemistry for Merchant Marine Oriented Individuals (Ch-1 and Ch-1C). A Study Guide to Minicourse D-1, Elementary Background.
PUB DATE Sep 78
NOTE 104p.; Contains occasional light and broken type

EDRS PRICE MF-$0.83 HC-$6.01 Plus Postage.
DESCRIPTORS *Chemistry; Course Content; Curriculum; *Higher Education; *Instructional Materials; Laboratory Procedures; *Oceanology; Science Education; *Short Courses; Study Guides

ABSTRACT

This document contains lab activities, problem sets, and a tape script to be accompanied by a slide show. The minicourse covers the following topics of general chemistry: kinetic-molecular theory, the Bohr atom, acids, bases, and salts, the periodic table, bonding, chemical equations, the metric system, computation of density, mass, and volume, and identification of various pieces of glassware used in a general chemistry laboratory. (BB)

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AN INTRODUCTION TO CHEMISTRY

FOR

MERCHANT MARINE ORIENTED INDIVIDUALS (Ch-1 AND Ch-1C)

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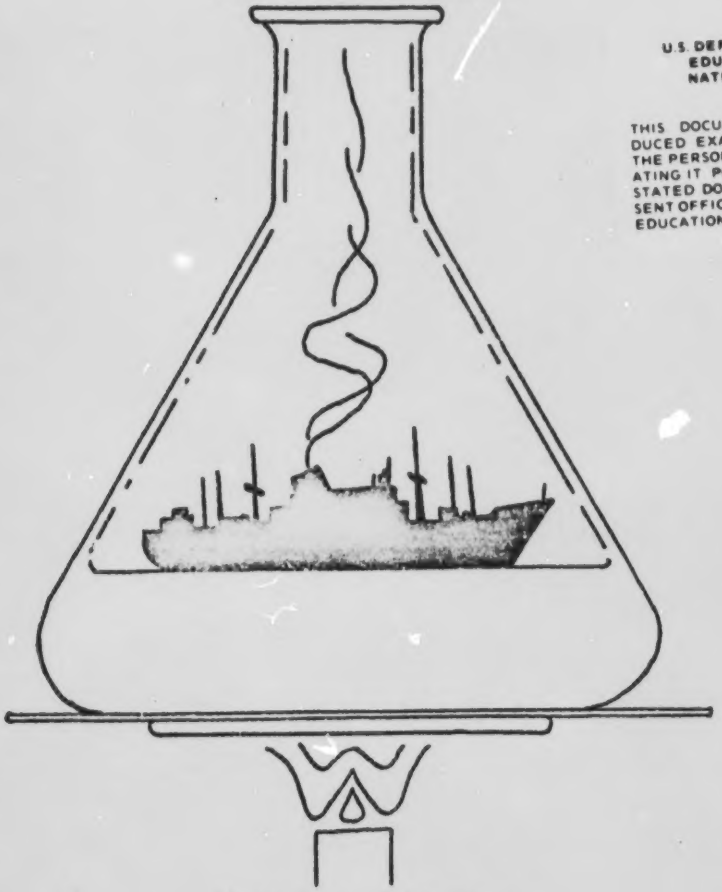
# A STUDY GUIDE TO

MINICOURSE D-1

ELEMENTARY BACKGROUND



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Maine Maritime Academy
Castine, Maine
September 1978



SE 026 370

## INTRODUCTION

Minicourse D-1 is the first of several minicourses you will be required to study as you progress through, *An Introduction to Chemistry for Merchant Marine Oriented Individuals*. It is intended to introduce students to the basic concepts of chemistry. For those students who have already studied some chemistry, this minicourse should prove a refreshing review. You should now obtain a D-1 slide carousel and D-1 audio tape or tape script and set up your equipment.

## MINICOURSE OBJECTIVES

As a result of studying this minicourse, the student should be able to:

1. Define phase, density, solubility, luster, brittleness, homogeneous, molecule, solvent, solute, element, symbol, acid, base, salt, indicator, neutralization, organic compounds, colloidal suspension, Tyndall effect, atom, electron, proton, neutron, isotope, ion, electrovalent bond, ionic bond, covalent bond, ionization, dissociation, precipation, formula, equation, radical, mixtures, and compounds.

2. Discuss kinetic-molecular theory, the organization of the periodic table, a simplified Bohr model of the atom, and the law of definite proportions.

3. Compute density, mass, and volume when two of the variables are given.

4. Use the metric system of measurement to determine volumes.

5. Identify by name the various pieces of glassware used in a general chemistry laboratory, and describe the uses of those pieces.

6. Describe the laboratory activities conducted during the study of this minicourse.

Turn now to Appendix A of the study guide and conduct the first laboratory

activity listed there.

TEXT INFORMATION AND PROBLEM SET

Appendix B contains a sheet titled, "Minicourse D-1 Problem Set." This sheet lists pages in the student text which support the audio-tape information. Also, this sheet provides the list of questions that constitute the problem set. You should look at this sheet now and begin to read the textual material and work the associated problems as you continue through the minicourse. Look at Appendix B now, and when you are ready, return to the audio-tape (students enrolled in Ch-1C should be following the same guide but need not submit the problems).

TERMS

See if you can define the following list of terms. Once you have written your definitions, compare them with those on the audio tape:

1. Substance _____
2. Homogeneous _____
3. Mixture _____
4. Heterogenous _____
5. Substance _____

6. Phase _____
7. Density _____
8. Solvent _____
9. Solute _____
10. Solubility _____
11. Luster _____
12. Malleability _____
13. Ductility _____
14. Tenacity _____
15. Brittleness _____
16. Conductivity _____
17. Element _____
18. Atom _____

Stop here, you will return after viewing slide number thirty-four.

19. Symbol _____
20. Metal _____
21. Metaloid _____
22. Compound _____
23. Acid _____
24. Base _____
25. Salt _____
26. Organic Compound _____
27. Neutralization _____
28. Titration _____
29. Colloid _____
30. Suspension _____
31. Tyndall Effect _____

32. Periodic Table _____

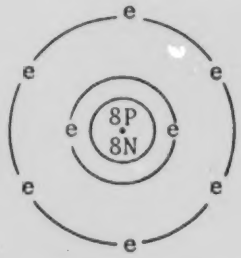
33. Element _____

Now return to the audio tape. Compare the answers you have written here with those given on the tape and in your student text. When you are satisfied, go on with the audio tape.

ORBITALS

1. The drawings you have just made of carbon and oxygen should look like those shown below.

CARBON




OXYGEN

Return to the audio tape.

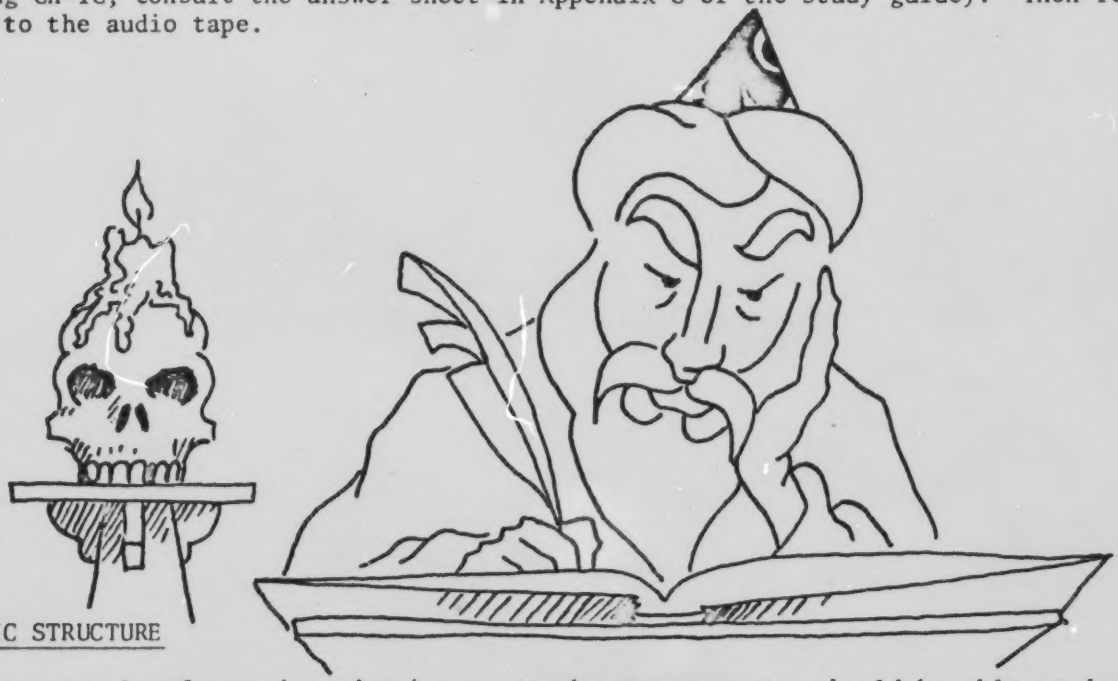
2. Complete the following chart for elements numbers 1-16:

| ELEMENT NUMBER | ELEMENT NAME | ELECTRONS IN FIRST ORBIT | ELECTRONS IN SECOND ORBIT | ELECTRONS IN THIRD ORBIT |
|---|---|---|---|---|
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |

| ELEMENT NUMBER | ELEMENT NAME | ELECTRONS IN FIRST ORBIT | ELECTRONS IN SECOND ORBIT | ELECTRONS IN THIRD ORBIT |
|---|---|---|---|---|
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |

a. Now draw a picture of each electron configuration on a separate sheet of paper.

b. Upon completion of this chart, show it to your instructor (if you are taking Ch-1C, consult the answer sheet in Appendix C of the study guide). Then return to the audio tape.



ATOMIC STRUCTURE

As a result of your introduction to atomic structure you should be able to define, illustrate, and answer the following as the case warrants:

1. Define the following:

   a. electron _____

   b. neutron _____

   c. proton _____

d. atomic mass units _____

e. positive ion _____

f. negative ion _____

2. Calculate in atomic mass units the atomic masses of the following:

a. HCL _____ d. NaCL _____

b. $H_2SO_4$ _____ e. $H_2O$ _____

c. NaOH _____ f. $N_2H_4$ _____

3. Calculate the numbers of electrons the following ions would have:

a. $H^+$ _____ d. $O^{-2}$ _____

b. $Na^+$ _____ e. $K^+$ _____

c. $CL^-$ _____

4. Some elements are classified as metals while others are classified as non-metals. On what bases are these classifications made? _____

_____

5. Draw the electronic structures of:

a. C d. Na

b. Ll e. AL

c. O f. N

When you have finished these questions, return to the audio tape.

VALENCE

The valence electrons are those in the outermost shell or orbital of an atom. How many valence electrons have each of the following atoms, and would you expect them to gain or lose electrons?

| | NUMBER | GAIN | LOSE |
|---|---|---|---|
| 1. H | _____ | ____ | ____ |
| 2. CL | _____ | ____ | ____ |

| | NUMBER | GAIN | LOSE |
|---|---|---|---|
| 3. O | _____ | ____ | ____ |
| 4. C | _____ | ____ | ____ |
| 5. Na | _____ | ____ | ____ |
| 6. Li | _____ | ____ | ____ |

Consult Appendix C for the correct answers, and then return to your audio tape.

BONDING

When molecules and their associated bonds are shown graphically, only the valence electrons are shown. Normally, X's are used to show the valence electrons of one element while dots are used to show the valence electrons of another element. Draw a picture each of the following molecules, showing their valence electrons:

1. NaCL
2. HCL
3. $O_2$
4. $CH_4$
5. $NH_3$
6. KCL

When you have finished, show your work to your instructor. If you are enrolled in Ch-1C consult Appendix C for the correct answers. Now return to the audio tape.

CHEMICAL EQUATIONS

In the following chemical equations, identify the formulae, radicals, and ions:

1. $H_2SO_4 \rightarrow 2H^+ + SO_4 =$
2. $C_8H_{18} + O_2 \rightarrow 8CO_2 + 9H_2O =$
3. $HCL + NaOH \rightarrow NaCL + H_2O =$
4. $HCL \rightarrow H^+ + CL^- =$

Consult Appendix C of this guide for the correct responses to this question.

This completes your work with minicourse D-1. If you are enrolled in Ch-1 and you have not already done so, remember that you must complete your D-1 problem set prior to taking the D-1 examination.

APPENDIX A

# LABORATORY ACTIVITIES

1. Lab Exercise 1, IDENTIFYING LABORATORY EQUIPMENT (Consult the lab equipment sheet if you are enrolled in Ch-1C, located in Appendix C of the study guide)

   A. Look at the glassware located on the shelf above your carrel. Locate the following: (1) 10 ml graduated cylinder; (2) 100 ml graduated cylinder; (3) an Erlenmeyer flask; (4) a 50 ml, 200 ml, and 5 ml beaker; (5) tongs; (6) medicine dropper; (7) watch glass; (8) burette; (9) motar and pestle; (10) evaporating dish; (11) Florence flask; (12) burette clamp; (13) ring stand; (14) test tube; (15) test tube holder; (16) extension clamp; (17) clay triangle; (18) test tube brush; (19) wash bottle; (20) funnel; (21) filter paper; (22) wire gauze; (23) stirring rod; (24) filter flask; (25) separating funnel; (26) triple arm balance; (27) analytical balance; and (28) stirring rod.

   B. You may want to make and name sketches of these items. You should be able to identify each with ease. Once you have completed your study of these items, turn on your audio tape or return to a copy of the tape script and begin.

2. Lab Exercise 2, DENSITY OF LIQUIDS.

   A. Determine the densities of the liquids in the bottles marked A, B, C, and D.

   B. Procedures:

      (1) Look at *slide number thirteen* and gather the equipment shown there.

      (2) Look at *slide number fourteen*. In reading a graduated cylinder correctly you must line up the bottom of the curved surface, called the meniscus, to obtain an accurate reading. The reading is made at the bottom of the meniscus.

      (3) Weigh the empty graduated cylinder and record its mass in each of the spaces provided.

      (4) For each of samples A, B, C, and D, pour a few ml in the cylinder, (one sample at a time) determine the mass of the combined cylinder, plus the sample, and record these combined masses in the spaces provided.

      (5) Record the volume of each fluid whose mass you have measured.

      (6) Determine mathematically the density of each fluid.

DATA

LIQUID A

1. Mass of graduated cylinder plus sample _____

2. Mass of graduated cylinder _____

3. Mass of sample _____

LIQUID A

4. Volume of sample _____ _____ Density of sample

LIQUID B

1. Mass of graduated cylinder plus sample _____
2. Mass of graduated cylinder _____
3. Mass of sample _____
4. Volume of sample _____ _____ Density of sample

LIQUID C

1. Mass of graduated cylinder plus sample _____
2. Mass of graduated cylinder _____
3. Mass of sample _____
4. Volume of sample _____ _____ Density of sample

LIQUID D

1. Mass of graduated cylinder plus sample _____
2. Mass of graduated cylinder _____
3. Mass of sample _____
4. Volume of sample _____ _____ Density of sample

(7) Answer the following questions:

a. What effect will changing the volume of a sample have upon the density? _____

b. What units of volume, mass, and density are used in scientific investigations? _____

(8) When you are satisfied with your work show it to your instructor. If it is satisfactory, he or she will give you credit for having completed the exercise.

(9) Proceed to lab exercise 3.

3. Lab Exercise 3, DETERMINING THE DENSITY OF SOLIDS

A. Determine the densities of solids A and B.

B. Procedures:

(1) Determine the mass of each solid and record this figure in the space provided.
(2) Determine the volume of each mass according to the formulas $v = \ell\omega h$ and $v = \pi r^2 h$ as the case warrants. Record these figures in the space provided.
(3) Mathematically compute the density in each case.

DATA

| SOLID A | SOLID B |
|---|---|
| (1) Mass of the solid _____ | (1) Mass of the solid _____ |
| (2) Volume of the solid _____ | (2) Volume of the solid _____ |
| (3) Density of the solid _____ | (3) Density of the solid _____ |

(4) When you have finished show your work to your instructor and he or she will give you credit for the exercise.

(5) If you are satisfied that you understand the concept of density, return to the audio tape. If you are not, and have already read the relevant sections in the student text, study the supplementary set of overhead projections on density. Your instructor will provide these upon request. (If you are enrolled in Ch-1C, these are located in Appendix C).

4. Lab Exercise 4, LAW OF DEFINITE COMPOSITION (ELECTROLYSIS OF WATER)

A. Procedures:

(1) Locate the equipment shown in slide number twenty-three, and insure that it is set up as shown.
(2) Measure the distance between the top of the liquid in the two outside tubes and their respective stoppers and record these distances.
(3) Sum these distances to give a total distance.
(4) Mathematically compute the percentage of the total that each individual distance is.
(5) Turn on the power supply and allow it to run for approximately 10 minutes, and then make a second set of calculations - the same as you did before the power supply was turned on.

DATA

| | BEFORE | | AFTER | |
|---|---|---|---|---|
| | Right Side | Left Side | Right Side | Left Side |
| Height to top | _____ | _____ | _____ | _____ |

| | BEFORE | | | AFTER | |
|---|---|---|---|---|---|
| | Right Side | Left Side | | Right Side | Left Side |
| Total | _____ | | Total | _____ | |
| % | _____ | _____ | % | _____ | _____ |

(6) Answer the following questions:

a. How do before and after percentages compare? _____

_____

b. How are your findings related to the law of definite proportions or constant composition? _____

_____

(7) When you have completed this activity, have your instructor check it and give you credit for having completed it. Then return to the audio tape.

5. Lab Exercise 5, ACID-BASE TITRATIONS, INDICATORS AND END POINTS

A. Look at slide *number twenty-six*. It shows some of the equipment and materials needed for this exercise. Gather them now! For the dilute acids ($H_2SO_4$) use .5M hydrochloric acid (Hcℓ) and .5M sulfuric acid and for the dilute base use .5M sodium hydroxide (NaOH). You will also need a ring stand, two burettes, and two extension clamps.

B. Procedures:

(1) Attach the extension clamp to the ring stand and a burette to each side of the extension clamp.
(2) Pour some Hcl in the left burette (as you face it) and some $H_2SO_4$ in the right burette.
(3) Pour 3 to 4 ml of NaOH in each of two test tubes. Stand one of the test tubes in a 250 ml beaker and 2 or 3 drops of phenolphthalien to the other.
(4) Record each burette reading.
(5) Hold the test tube with NaOH and phenolphthalien under the burette containing Hcl and add Hcl dropwise (shake the test tube while you are doing this) to the test tube until the color just changes. Record the new Hcl burette reading.
(6) Repeat the above process, using the other test tube and the $H_2SO_4$ burette.
(7) Empty the Hcl burette, rinse it with distilled water, pour some NaOH into it, and take a burette reading.
(8) Pour 3 to 4 ml Hcl into a test tube, and add 2 or 3 drops of phenolphthalien.
(9) While gently shaking the test tube, add NaOH dropwise until the color just begins to turn red, and record the final burette reading.

DATA

| #1 - Hcl | #2 - $H_2SO_4$ | #3 - NaOH |
|---|---|---|
| Initial burette reading _____ | Initial burette reading _____ | Initial burette reading _____ |
| Final burette reading _____ | Final burette reading _____ | Final burette reading _____ |
| Hcl used _____ | $H_2SO_4$ used _____ | NaOH used _____ |

(10) Answer the following questions:

a. What results were observed when you added phenolphthalein to the test tube containing NaOH? _____
_____.

b. What results were observed when you added phenolphthalein to Hcl? _____
_____.

c. What is your hypothesis concerning the color changes that were observed? _____
_____.

(11) Save these results, you will need to use them in a later mini-course.

(12) Return to the audio tape and listen to the point where the slide number twenty-seven discussion begins.

6. Lab Exercise 6, ACID-METAL REACTIVITY

A. The purpose of this exercise is to investigate the effect of an acid of varying strengths upon a metal.

B. Procedures:

(1) You will actually conduct this activity three times, once with .25M $H_2SO_4$, once with 1.0M $H_2SO_4$, and once with 2.0M $H_2SO_4$.

(2) Add two or three zinc chips to a test tube and then cover the zinc with .25M $H_2SO_4$.

(3) Hold the test tube at eye level and observe. Record your observations in the space provided.

(4) Repeat the preceeding operation, using 1.0M and 2.0M $H_2SO_4$. Record your observations in the spaces provided.

OBSERVATIONS

1. .25M $H_2SO_4$ + Zn _____

_____

2. 1.0M $H_2SO_4$ + Zn _____

_____

3. 2.0M $H_2SO_4$ + Zn _____

_____

(5) Answer the following questions:

a. What happened as the strength (.25M is not as strong as 2.0M) of the $H_2SO_4$ increased? _____

_____

b. What were the bubbles you observed to rise off of the zinc? _____

_____

c. What happens to the zinc as a result of this mixing of $H_2SO_4$ + Zn? _____

_____

(6) Discard the acid you have used and place your zinc pieces in the waste zinc container.

(7) Return to the audio tape.

7. Lab Exercise 7, TITRATION OF SALINE WATER

A. The purpose of this exercise is to give the student experience titrating saline water.

B. You will use the following equipment in conducting this exercise; local sea water (sample supplied for Ch-1C students), N/10 silver nitrate (AgNo3), a burette and associated equipment, potassium chromate indicator ($K_2CrO_4$), 250 ml. Erlenmeyer flash test tube, pipet, filter paper, 100 ml. graduated cylinder, distilled water.

C. Procedures:

(1) Filter 10 ml. sea water into the 250 ml. Erlenmeyer flask.
(2) Add 90 ml. distilled water to the flask.

(3) Add 5 drops $K_2CrO_4$ to the flask.
(4) Swirl flask until contents are thoroughly mixed.
(5) Fill the burette with approximately 50 ml. $AgNO_3$.
(6) Record the burette reading to the nearest 0.01 ml.
(7) Add $AgNO_3$ dropwise to the flask while constantly swirling the flask. Continue process until one drop produces a reddish color that is maintained for 30-40 seconds (this is the end point).
(8) Record burette reading to nearest 0.01 ml.
(9) Repeat the process, using a second seawater sample.

| DATA | RUN 1 | RUN 2 |
|---|---|---|
| a. $AgNO_3$ initial volume | _____ | _____ |
| b. $AgNO_3$ final volume | _____ | _____ |
| c. $AgNO_3$ used | _____ | _____ |
| d. Volume of sample | _____ | _____ |

(Save these data, you will use them later.)

D. The method you have just used is known as the Mohr method of chloride analysis. With this method, the chloride ion is precipated as silver chloride (AgCl). Once all of the chloride has been precipated, the excess silver ion forms a red, silver chromate precipate ($Ag_2CrO_4$). The two reactions involved in this process are:

(1) $Ag^+ + Cl^- \rightarrow AgCl$.

(2) $2Ag^+ + CrO_4^{2-} \rightarrow AgCrO_4$.

E. Once you have completed the activity, show your data to your instructor. When you have received credit for the exercise, return to the audio tape.

APPENDIX B

# PROBLEM SET

# MINICOURSE D-1 - PROBLEM SET

## INTRODUCTION

The problem set listed below was taken from your course text, Brown, T. H., and LeMay, E. H., *Chemistry - The Central Science*, Prentice-Hall, Inc., Englewood Cliffs, New Jersey, 1977. It must be completed with a grade of 70% prior to taking the D-1 Minicourse examination for the first time. You may submit the problem set as many times as are necessary in order to attain the 70% level.

The following sections in your text are relevant to Minicourse D-1 as well as to the completion of the problem set. Those sections are Chapters 1 and 2, pages 84-85 (titration), and pages 472-477 (indicators). In addition, Chapters 1 and 2 of Hill, J. C., *Student's Guide to Brown and LeMay, Chemistry - The Central Science*, are also applicable.

## PROBLEM SET

Before beginning to work the problems listed below, look closely at the exercises at the backs of the chapters and then at Appendix F in the back of the text. Note in the exercise sections of the chapters that selected problems have their item numbers printed in blue. The answers to these exercises are located in Appendix F. For every problem whose answer is given, there is usually a companion problem whose answer is not given. You might use this as a guide; if you can correctly solve the problem whose answer is given, you can probably solve the one whose answer is not given.

The following questions should be answered and submitted to your instructor for evaluation. They should be neatly written and each successive step in the solution should be shown. (Read across)

| PAGE | PROBLEMS |
|---|---|
| 23 | 1.8, 1.10, 1.12 |
| 24 | 1.34, 1.36, 1.40, 1.42 |
| 56 | 2.2, 2.8, 2.12, 2.14, 2.24, 2.26 |

APPENDIX C

# SUPPLEMENTARY EXERCISES AND ANSWERS TO STUDY GUIDE ACTIVITIES

1. Laboratory glassware can be found in the pamphlet titled Laboratory Equipment, Glassware, Procedures, and Safety Precautions.

2. CH-1C students will find the supplemental density and atomic structure sheets in the packet marked Density and Atomic Structure.

DENSITY

The first step in solving these types of problems is to draw a picture.

Here we are dealing with the following relationship:

$$\text{DENSITY} = \frac{\text{MASS}}{\text{VOLUME}} \text{ OR } D = \frac{M}{V}$$

From this we see that the density of an object is directly proportional to the object's mass and inversely proportional to its volume. The basic formula can be manipulated as follows:

(a) $\frac{V}{1} \times D = \frac{M}{V} \times \frac{V}{1}$

(b) $M = DV$

and

(a) $\frac{1}{D} \times M = DV \times \frac{1}{D}$

(b) $V = \frac{M}{D}$

1. What is the density of a cube whose dimension along one side is 10 mm and whose mass is 100 g? The answer should be in grams per cubic centimeter (grams/cubic centimeter, g/cc, or $g/cm^3$).

   (a) First draw and label the picture.

   (b) Convert millimeters to centimeters.

   (c) Compute the volume of the cube using the formula:

   Volume = length x width x height

   V = 1 cm x 1 cm x 1 cm

   $V = 1\ cm^3$

   (d) Solve, using the formula: D = M/V

   $D = 100g/1\ cm^3$

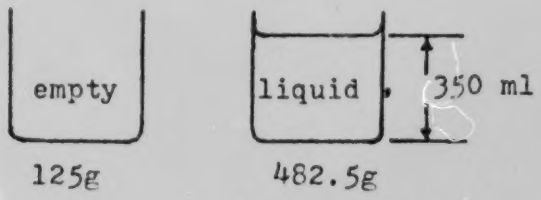
   $D = 100g/cm^3$

   The answer means that each cubic centimeter of this material has a density such that its mass is 100g.

2. What is the density of a clear, watery sample of liquid if the mass of its con-

tainer, empty, is 125g, the mass of the container plus the liquid is 482.5g and the measured volume is 350 ml? Finally, could the sample possibly be seawater?

a.




b. Compute the mass of the liquid alone;

482.5g - 125g = 357.5g

c. Compute the density of the liquid by using the formula;

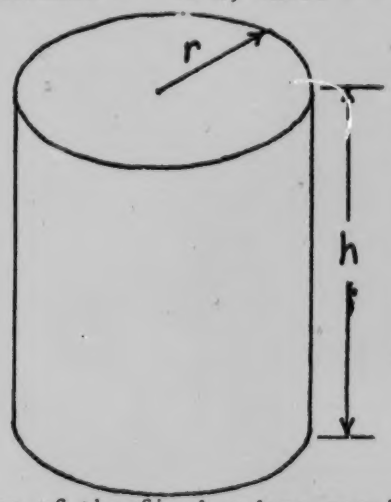
D = M/V  Assume 1 ml = 1cc

D = 482.5g/357.5cc

D = 1.03g/cc

d. Look up the density of seawater. The density of seawater at a salinity of 30 parts per thousand (0/00) is 1.025. The answer then to the last question is "Yes, the sample could be seawater".

3. When dealing with density problems which require the computation of the volume of a cylinder, use the formula $V = \pi r^2 h$, where $\pi = 3.1416$, r = radius, and h = height.




Remember that computation of the final volume requires use of the inside radius of the container, and not the outside radius.

ORBITALS

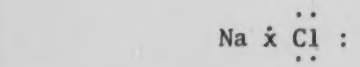
| ELEMENT NUMBER | ELEMENT NAME | NO. ELECTRONS IN FIRST ORBITAL | NO. ELECTRONS IN 2ND ORBIT | NO. ELECTRONS IN 3RD ORBIT |
|---|---|---|---|---|
| 1 | HYDROGEN | 1 | | |
| 2 | HELIUM | 2 | | |
| 3 | LITHIUM | 2 | 1 | |
| 4 | BERYLLIUM | 2 | 2 | |
| 5 | BORON | 2 | 3 | |
| 6 | CARBON | 2 | 4 | |
| 7 | NITROGEN | 2 | 5 | |
| 8 | OXYGEN | 2 | 6 | |
| 9 | FLUORINE | 2 | 7 | |
| 10 | NEON | 2 | 8 | |
| 11 | SODIUM | 2 | 8 | 1 |
| 12 | MAGNESIUM | 2 | 8 | 2 |
| 13 | ALUMINUM | 2 | 8 | 3 |
| 14 | SILICON | 2 | 8 | 4 |
| 15 | PHOSPHORUS | 2 | 8 | 5 |
| 16 | SULFUR | 2 | 8 | 6 |

VALENCE

| | NUMBER | GAIN | LOSS |
|---|---|---|---|
| 1. H | 1 | | X |
| 2. Cl | 7 | X | |
| 3. O | 6 | X | |
| 4. C | 4 | NEITHER GAIN OR LOSS | |
| 5. Na | 1 | | X |
| 6. Li | 1 | | X |

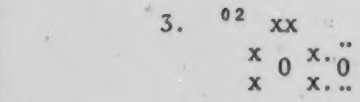
BONDING

1. NaCl



2. HCl

3. $O_2$



4. $CH_4$



5. $NH_3$



6. KCl

CHEMICAL REACTIONS

1. $H_2SO_4 \rightarrow 2H^+ + SO_4 =$

   Formula Ion Radical

2. $C_8H_{18} + O_2 \rightarrow 8CO_2 + 9H_2O$

   Formula Formula Formula Formula

*3. $HCl + NaOH \rightarrow NaCl + H_2O$

   Formula Formula Formula Formula

4. $HCl \rightarrow H^+ + Cl^-$

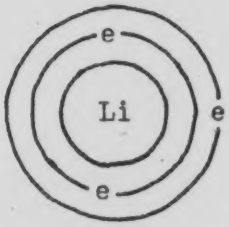
   Formula Ion Ion

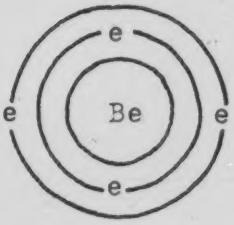
* An example of an acid base neutralization and the resultant formation of a salt and water.

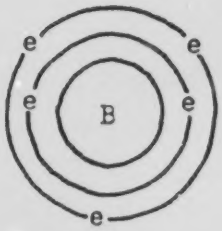
ATOMIC STRUCTURE

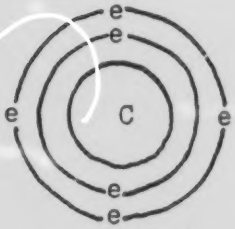
1.

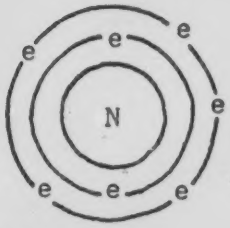

2.

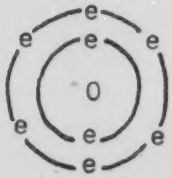

3. 


4. 


5. 


6. 


7. 


8. 

APPENDIX D

# TAPE SCRIPT

# TAPE SCRIPT FOR MINICOURSE D-1

## ELEMENTARY BACKGROUND

| SCRIPT | SLIDES AND NOTES |
|---|---|

MINICOURSE D-1

ELEMENTARY SUBSTANCES

Advance your slide carousel to *Slide #1*. Three definitions are presented on this slide.

Substances are chemicals whose properties and composition are identical throughout a given sample. Water ($H_2O$) is an example of a substance. No matter what the source, pure water is always the same.

Substances are homogeneous, that is, they are the same throughout. Let us consider pure water again. All portions of pure water ($H_2O$) are the same.

A mixture is composed of two or more distinct substances. These substances differ in their properties and composition. A mixture may not be homogeneous throughout. For example, sand and water ($H_2O$) form a

Substances include a variety of matter, all specimens of which have identical properties and composition.

Substances are homogeneous.

A mixture consists of two or more distinct substances differing in properties and composition.

SLIDE #1

| SCRIPT | SLIDES AND NOTES |
|---|---|

heterogeneous mixture. That is, they may be separated into their component parts. However, salt (NaCl) and water ($H_2O$) form a homogeneous mixture since they cannot be readily separated.

Look at *Slide #2*. Salt (NaCl) is a pure substance consisting of sodium (Na) and chlorine (Cl) chemically bonded together. All samples of salt (NaCl) will have identical compositions. Salt (NaCl) and pepper on the other hand form a mixture. The substances are distinct entities as can be seen in the slide. The pepper may or may not be evenly distributed in the salt (NaCl). That is, the mixture may or may not be homogeneous. However, salt (NaCl) is always homogeneous.

Advance your carousel to *Slide #3*. The mixture you see is sand and iron filings. The iron filings may be separated from the sand because they are magnetic, while the sand is not. Generally, mixtures may be separated by physical means. The components of substances cannot be separated physically.

SALT AND PEPPER with a pile of salt in front of the salt shaker and a pile of pepper in front of the pepper shaker.

SLIDE #2

Two piles of sand and iron filings mixed with a magnet held close to one pile.

SLIDE #3

SCRIPT

Look at *Slide #4*. Materials dissolved in water constitute a mixture which can be separated by boiling. If, for example, steam can escape from a boiler, soon all that is left is a mineral residue (minerals and water form a mixture).

Look at *Slide #5*. So far we have talked about substances and mixtures. Can you write their definitions? See if you can and then compare your definitions with those on Slide #1.

Now, advance your carousel to *Slide #6* and let us consider some physical properties of matter. The term "phase" is used to refer to the form of matter. Matter may exist as a solid, a liquid, or a gas. Gases are a form of matter which has neither definite form nor definite shape. Liquids have definite volume but not definite shape. Solids have both definite volume and definite shape.

Look at *Slide #7*. The drop of water ($H_2O$) you see isn't very large. But consider how much smaller it could

SLIDES AND NOTES

Two 250 ml. beakers. One beaker contains a solution while the other beaker contains the solute after the water has been boiled off.

SLIDE #4

SUBSTANCE

HOMOGENEOUS

MIXTURE

SLIDE #5

The term "phase" is used to refer to the gas, liquid, or solid form of matter.

Gases have neither definite shape or volume. They assume the shape and volume of their container.

Liquids have a definite volume, but take the shape of their container.

Solids have definite shape and volume.

SLIDE #6

SCRIPT

be made. It could be cut in half, and one of the halves cut in half; and so on, seemingly to infinity. But can this go on to infinity? Is there a limit to how small a drop of water ($H_2O$) can be, and still be a drop of water? If your answer was "Yes", you were correct.

The ancient Greeks considered this question and decided there was a limit. They then coined a word for this limit. They called the ultimately small particle an atom. Today, we know that the smallest particles of everything but elements can be broken down to elements. For most substances, however, the limiting particle is called a molecule. Water ($H_2O$) then can be divided only to the point where there is one molecule. At this point, further division requires that the molecule be broken into its component hydrogen (H) and oxygen (O) atoms.

Look at *Slide #8*. A molecule can be thought of as being the smallest particle of a substance that still has

SLIDES AND NOTES

SEVERAL WATER DROPLETS ON A LEAF

SLIDE #7

A molecule can be thought of as being the smallest particle of a substance that still has the properties of that substance.

SLIDE #8

SCRIPT

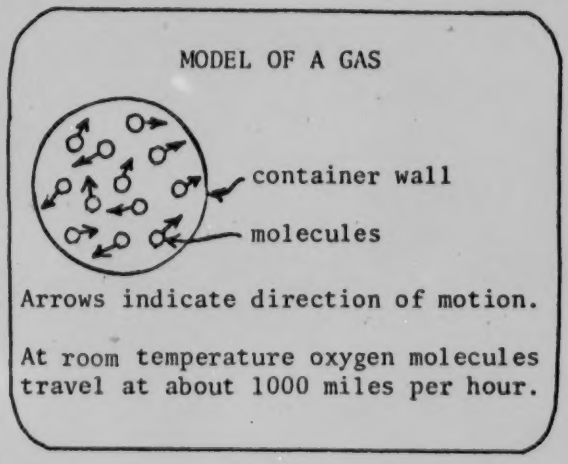
SLIDES AND NOTES

the properties of that substance.

Now advance your carousel to *Slide #9*. The state of matter (solid, liquid, or gas) depends on the energy level of the molecules. In the gaseous state the molecules have sufficient energy to break free of each other and if they collide, the collisions are violent enough to prevent their sticking together.

When molecules in the gaseous state bounce off the walls of their container, the force of the collisions with the walls causes what we call pressure. The speed of the molecules upon impact with a container's walls is proportional to the force which they create. For example, at room temperature oxygen ($O_2$) molecules have a speed of 500 meters per second, or 1,000 miles per hour.

In the liquid and solid states, atoms or molecules do not have enough energy to break the forces of attraction between them. In the solid phase they line up geometrically much as


MODEL OF A GAS
container wall
molecules
Arrows indicate direction of motion.
At room temperature oxygen molecules travel at about 1000 miles per hour.

SLIDE #9

SCRIPT

SLIDES AND NOTES

spheres might be expected to do. Note that molecules are stationary in the solid phase; however, this does not mean that they do not have motion. On the contrary, they vibrate and rotate in place.

Look at *Slide 10*. Here you will find an illustration of the concept about which we have just talked. In the liquid phase atoms or molecules, as the case may be, have enough energy to break out of the geometric pattern but not enough to separate.

Not all molecules have identical energies at any given time. Some may be stopped while others move very rapidly. This accounts for the phenomenon called evaporation.

Now advance your slide carousel to *Slide #11*. If a surface molecule acquires sufficient energy to break the forces holding it to its neighbors it might "pop" out of the surface and become a free gas molecule. It acquires the energy to do this from collisions with other fast moving

Three cubic containers: one showing particles in the gaseous state, one showing particles in the liquid state, and one showing the arrangement of particles in the solid state.

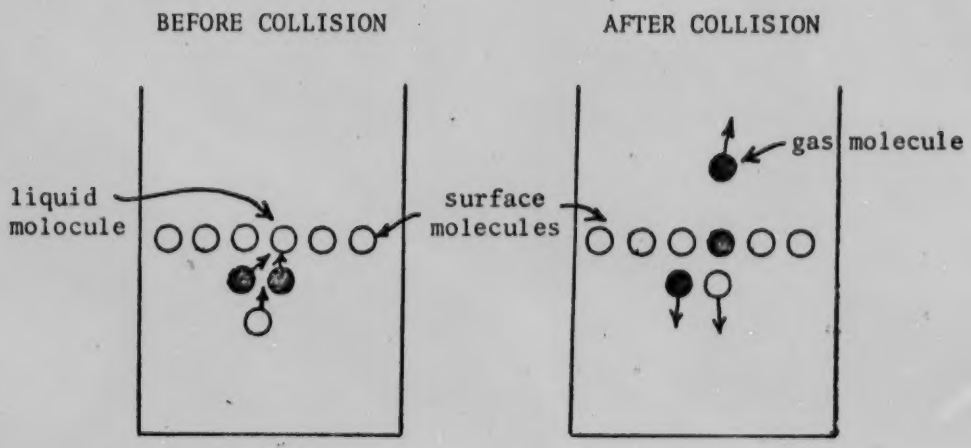
SLIDE #10

*(SEE PAGE 31A -FOLLOWING)*

SLIDE #11

MODEL FOR EVAPORATION

SLIDE #11




If a surface molecule (red) gets hit just right, it might "pop" out of the surface to become a gas molecule.

SCRIPT

molecules that hit it from below.

Look at *Slide #12*. Density is another term frequently used to describe matter. Density is the ratio of mass to volume, or D = M/V. Sometimes density is defined as the weight to volume ratio. These are the definitions that you must know in order to do your calculations.

Stop your tape at this point and read the section in your minicourse study guide titled "Text Information and Problem Set."

At this point, you will conduct two density experiments, one dealing with liquids and the other dealing with solids. These are listed as lab exercises 2 and 3 in Appendix A. Conduct them now. *(Slides #'s 13 and 14* are used with lab exercises 2 and 3).

Advance your slide carousel to *Slide #15*. Let us consider solutions. Slide #15 lists some basic terms related to solutions. A solvent is a material in which another dissolves. The solute is the dissolved material.

SLIDES AND NOTES

Density = $\frac{\text{mass}}{\text{volume}}$ D = M/V or m/v

For example:

D = grams/cubic centimeter or g/cc
D = kilograms/liter or kg/l
D = grams/milliliter - g/ml

Sometimes weight density:

D = pounds/cubic foot = lb/cuft
= lbs/gallon

SLIDE #12

A triple arm balance, graduated cylinder, and four 250 ml. sample bottles marked "A", "B", "C", and "D". The sample bottles contain four different solutions (different colored solutions).

SLIDE #13

A gradutated cylinder showing a meniscus at 96.4 ml.

SLIDE #14

SCRIPT

Solubility is the ability, or a measure of the ability, of one material to dissolve in another. For example, if sugar ($C_6H_{12}O_6$) dissolves in water ($H_2O$), it is said to be soluble. The sugar ($C_6H_{12}O_6$) is the solute; the water is the solvent.

Look at *Slide #16*, and let's consider some other descriptive terms. Luster is the silvery shine of many materials, particularly metals.

Look at *Slide #17*. Malleability is the ability to be pounded into thin sheets. It is the opposite of brittleness.

Look at *Slide #18*. Ductility is the ability to be drawn into a wire. Tenacity or tensile strength is the ability to resist being pulled apart. Brittleness is the tendency to shatter when struck with a sharp blow. Electrical conductivity is the ability of a substance to pass an electric current.

These are basic terms used to describe matter. You should be able to

SLIDES AND NOTES

Define: SOLVENT

SOLUTE

SOLUBILITY

SLIDE #15

Luster is the silvery shine of many materials - particularly metals.

Slide shows a shiny brass and steel padlock.

SLIDE #16

Malleability is the ability to be pounded into thin sheets.

Slide shows a piece of metal drawn into a long, thin shape.

SLIDE #17

SCRIPT

SLIDES AND NOTES

recognize and define them. Can you? Stop the tape and go to the section of the minicourse guide titled "Terms."

The question to be answered now is how can matter be classified? The usual scheme used by chemists is to classify matter as elements, compounds, and mixtures. We'll consider them one at a time.

Look at *Slide #19*. An element is a substance which cannot be decomposed by chemical means. A key word here is "chemical means." Elements can be changed by thermonuclear reactions or fission reactions.

Advance your carousel to *Slide #20*. The symbols on the chart represent the name of an element. In the upper right hand corner "H" stands for hydrogen; "Li" below it for lithium, and so on. The arrangement of the chart is by increasing atomic number. Generally, this means that as the mass of the atom increases, the atomic number increases. Hydrogen (H) is the lightest element known, Helium (He) is

Ductility is the ability to be drawn into a wire.

Tenacity, or tensile strength, is the ability to resist being pulled apart.

Brittleness is the tendency to shatter when struck with a sharp blow.

Electrical conductivity is the ability of a substance to pass an electric current.

SLIDE #18

An element is a substance which cannot be decomposed by chemical means.

Metals have luster, are good conductors of heat and electricity and are malleable and ductile.

Nonmetals are elements which have no luster, are not good conductors of heat and electricity, and when solid, are brittle.

Metaloid are elements which have some properties of both metals and nonmetals.

SLIDE #19

THE PERIODIC TABLE

SLIDE #20

SCRIPT

SLIDES AND NOTES

the next lightest, and so on. The chart is so arranged so that verticle columns contain elements that are chemically similar. In the first column, Lithium (Li), sodium (Na), Potassium (K), and rhubidium (Rd), for example, react with other elements in a very similar fashion.

Notice the heavy step-wise line near the middle of the chart. This line separates the metals from the non-metals. The metals are to the left of the line and the non-metals to the right.

Generally speaking, as one proceeds across a horizontal row or period, the elements become increasingly nonmetallic in character. The last column starting with Helium at the top, is the inert element column. This column could also be considered the divider between metals and non-metals.

The majority of elements are metals; they have lustre, are good conductors of heat and electricity, are malleable and ductile. All metals

SCRIPT

except mercury are solids at room temperature.

All non-metals, except bromine, are gases or brittle solids: they have no luster and are poor conductors of heat and electricity.

The elements along the zig-zag line but to the right of it, are called metaloids. These have properties of both metals and non-metals. Sometimes these properties are strange. Silicon, for example, under the right conditions, conducts electricity, but only in one direction.

Advance your slide carousel to *Slide #21*. This slide presents some terms and concepts with which beginning chemistry students should be familiar.

Look at *Slide #22*. A compound is a substance which can be decomposed by chemical means. It is composed of 2 or more elements. All samples of a compound have the same composition. Furthermore, the elements which are present in a compound are present in a

SLIDES AND NOTES

ELEMENT
METAL
NONMETAL
METALOID

Numerical arrangement of the periodic table of Elements.

Verticle arrangement of the periodic table.

Horizontal arrangement of the periodic table.

SLIDE #21

A compound is a substance which can be decomposed by chemical means.

A compound is made up of two or more elements.

All samples of a compound have the same composition.

The elements in a compound are present in a definite mass ratio which cannot be varied.

SLIDE #22

SCRIPT

definite ratio which cannot be varied.

Stop the tape and conduct laboratory activity #4. Its directions are located in Appendix A of the minicourse study guide.

You have just investigated the ratio of hydrogen (H) to oxygen (O) in water ($H_2O$). You should have found the ratio of one to the other the same before as after the experiment. This ratio is invariant, a demonstration of the law of constant or definite proportions. The law suggests that the ratio of elements in a molecule of a substance are in definite proportions to one another and that the proportion or ratio cannot or does not change.

Now look at *Slide #24*. Let us consider the classification of compounds. One way to classify them as acids bases, salts, and organic compounds.

Acids have a sour taste, for example, lemon is acid.

Acids affect indicators. Phenolphthalien turns colorless in acid. Tea becomes lighter when lemon is added.

SLIDES AND NOTES

Slide shows ELECTROLYSIS EQUIPMENT

SLIDE #23

Acids are a class of compounds which:

- have a sour taste
- affect the color indicators
- neutralize bases
- react with metals to produce hydrogen
- react with carbonates to liberate carbondioxide

SLIDE #24

SCRIPT

Acids neutralize bases, that is, acids counteract the effect of bases, and vice versa.

Acids react with metals to produce hydrogen, and of course, destroy the metal.

Acids react with carbonates to liberate $CO_2$. This property is often used in minerology to identify carbonate minerals.

Advance your carousel to *Slide #25*.

Return to Appendix A and conduct laboratory #5.

You have just conducted an acid base neutralization.

At the point where the color appeared and disappeared, the acid and base had neutralized each other and the solution was neither acid nor base but rather a salt. The process of finding the balance (neutral) point, so to speak, between acid and base is called titration. The titration process is used extensively in chemistry. For example, the process is used to

SLIDES AND NOTES

WARNING

STRONG ACIDS AND BASES DESTROY LIVING TISSUE. HANDLE WITH CARE!

If you come in contact with strong reagents, flush with water immediately.

SLIDE #25

Slide shows: two liter bottles, one containing dilute acid and one dilute base; a small bottle of phenolphthalein; two 50 ml. beakers; one 250 ml. beaker; one test tube; and two eyedroppers.

SLIDE #26

SCRIPT

SLIDES AND NOTES

analyze sea water, and especially to determine the salinity of a sample of sea water. In this case silver nitrate ($AgNO_3$) of known strength is used with sea water whose salinity is to be determined.

Titration is also used extensively in marine boiler water analysis to determine the constituents of the water. This type of analysis, along with others, uses some of the reagents you have just used in your laboratory activity.

If you are sure at this point that you understand the activity you have just conducted, see your instructor and have him or her credit you with completion of this activity. In the event that you still have some questions, consult the problem set sheet for this minicourse. It lists a section in your student text which discusses indicators. Do these things now.

Look at *Slide #27*. The slide shows the materials necessary to con-

Slide shows: one bottle containing dilute $H_2SO_4$; a few pieces of zinc; one 250 ml. beaker; and one test tube.

SLIDE #27

SCRIPT

SLIDES AND NOTES

duct laboratory activity #6. Collect these materials and then go to Appendix A, laboratory activity #6. Conduct this activity now.

When you poured $H_2SO_4$ over the mossy zinc (Zn) you should have observed bubbles to be rising from the zinc (Zn). These were bubbles of hydrogen (H) gas. As the strength of the acid was increased, you should have observed that the rate at which the bubbles formed also increased.

Hydrogen gas is explosive. Since this is the case, you should be able to pour some 2M $H_2SO_4$ over a couple of Zinc (Zn) pieces, hold a finger over the end of the test tube for a few seconds, and then as you move your finger away from the top of the tube, and bring a lighted match close to the opening, observe a small explosion. Try it if you would like.

The reaction you have just observed causes a deterioration in the zinc (Zn) metal.

$Zn + H_2SO_4 \rightarrow H_2(g) + ZnSO_4$

SCRIPT

Take your lab sheets to your instructor and have him credit you with completion of the activity.

Look at *Slide #28*, and gather the materials shown in that slide.

Place a marble chip in the test tube and add some acid (use 1.0 M $H_2SO_4$). Try to ignite the gas that comes off. Is it hydrogen?

When you have finished your observations, clean and return all of your materials.

Advance your carousel to *Slide #29*. A salt can be thoguht of as the product of the neutralization of any acid and any base. An acid and a base always combine to form a salt and water ($H_2O$). Most salts are found in minerals of the earth and in sea water. Table salt is the most common example.

Look at *Slide #30*. Organic compounds are compounds containing carbon (C). Methane ($CH_4$), or natural gas, consists of a carbon (C) atom with 4 carbons attached. Octane ($C_8H_{18}$) in gasoline has the configuration which

SLIDES AND NOTES

Slide shows: one liter bottle containing dilute acid; a bottle of marble chips; and a test tube.

SLIDE #28

*(SEE PAGE 42A - FOLLOWING)*

SLIDE #29

*(SEE PAGE 42B - FOLLOWING)*

SLIDE #30

| SCRIPT | SLIDES AND NOTES |
| --- | --- |
| is shown. Sugar also is a carbon (C) compound.<br><br>Carbon (C) atoms have the unusual property of being able to attach to each other to form long chains. This chain structure is a common characteristic of organic compounds.<br><br>Some compounds of carbon (C) are not considered to be organic. For example, limestone, which is calcium carbonate, is not organic. There are, however, not many of these compounds. | |

MODULE D1

SLIDE #29

A salt, which is a combination of a metal and a nonmetal, can also be considered to be the product of the neutralization of any acid and any base.

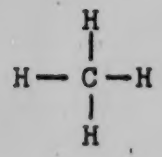
ACID + BASE → SALT + WATER

Hydrochloric acid + sodium hydroxide → sodium chloride + water

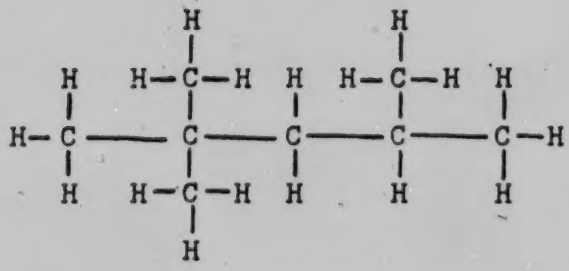
Acetic acid + ammonium hydroxide → ammonium acetate + water

MODULE D1

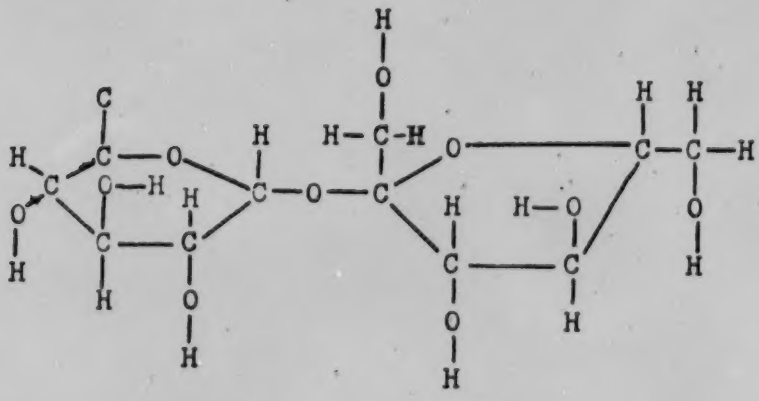
SLIDE #30



Methane



Octane



Ordinary Sugar

SCRIPT

Look at *Slide #31*. A mixture, you will recall, consists of 2 or more distinct substances. A suspension is a mixture containing particles of one substance distributed through another. Particles in suspension usually settle out. The muddy water in Slide #31 will eventually clear. Most suspensions are of this kind.

Look at *Slide #32*. Some suspensions, however, do not settle out. If the particles in suspension are small enough, the suspension is stable. Smoke is a common example of this type of suspension. Suspensions of this kind are called colloids. Colloidal suspensions may occur in the liquid or gas phase. Colloidal suspensions,or dispersions, as they are sometimes called, have the ability to scatter light.

Look at *Slide #33*. The beam of light is visable as it passes through the suspension.

This scattering of light by a colloidal dispersion is called the

SLIDES AND NOTES

Slide shows: a 250 ml. beaker containing muddy water.

SLIDE #31

Slide shows: a smoke stack emitting smoke into the environment.

SLIDE #32

Slide shows: a 500 ml. beaker containing coffee. This is a demonstration of the Tyndall effect (shows a light beam passing through the colloidal suspension).

SLIDE #33

SCRIPT

Tyndall Effect. It is sometimes used to measure concentrations of particles.

If the colloidal particles become so small that they are molecules or ions, the suspension becomes a solution. Solutions are dispersions of molecules or ions of one substance through another.

Look at *Slide #34*. Can you define these terms? Stop the tape and go to the terms section of the study guide.

In the next section we will consider a simplified model of the structure of matter. You will recall that matter cannot be divided into smaller and smaller parts indefinitely.

When molecules are chemically divided, atoms are produced, and atoms are the stuff of which matter is made. There are 92 naturally occurring species of atoms which we call elements.

Look at *Slide #35*. An atom is the smallest particle of an element that can exist separately or enter into combination with other atoms. All

SLIDES AND NOTES

Element - definition and characteristics.
Symbol.
Layout of the Periodic Chart (Table).
Metal.
Nonmetal.
Metaloid.
Compound - definition and characteristics.
Characteristics of: acids, bases, salts, organics.
Neutralization.
Suspension.
Colloid. Tyndall effect.

SLIDE #34

An atom is the smallest particle of an element that can exist or enter into combination with other atoms.

SLIDE #35

SCRIPT

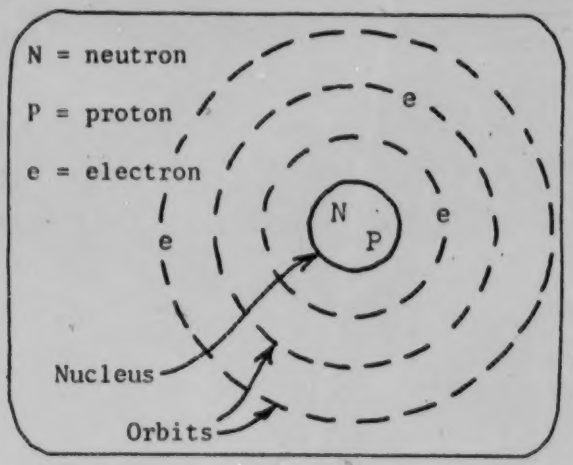
SLIDES AND NOTES

atoms of the same element are chemically alike.

Atoms can be broken down; however, not by chemical means. When they are broken down, all atoms are made up of 3 basic particles: electrons, protons, and neutrons. Atoms of different elements simply contain different members of these basic particles.

Advance your slide carousel to *Slide #36*. It is thought that the basic particles are arranged in an atom as depicted in the diagram. Protons and neutrons comprise a heavy nucleus that is positively charged. The nucleus is very small; about one trillionth of a centimeter in diameter. Electrons are smaller and are about 10,000 radii of the nucleus out from the nucleus. Note also that the electrons are thought to be arranged in orbits.

As an analogy, if the nucleus of an atom were the size of a basketball, the nearest electron would be about two miles away and substantially




SLIDE #36

SCRIPT

SLIDES AND NOTES

smaller than the basketball. Atoms consist primarily of empty space with an incredibly dense nucleus; about a trillion times as dense as water.

Look at *Slide #37*. The normal set of weights and measures is not practical when working with atoms.

Protons have a weight of 1 atomic mass unit, or amu, and an electrical charge of +1.

Neutrons weigh 1 amu and have no charge.

Electrons weigh practically nothing and have a -1 charge.

This difference in charge holds the atom together.

Look at *Slide #38*. The motion of the electrons keeps them from falling into the nucleus, much as the earth's motion keeps it from falling into the sun.

Remember from your study of physics that two forces come into play here. They are centrifugal and centripetal forces.

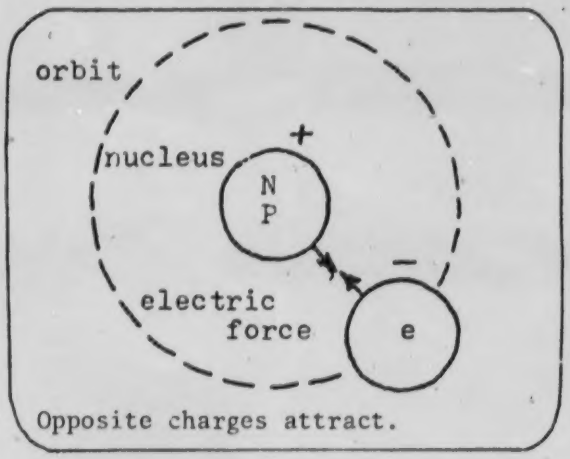
Look at *Slide #39*. What is the difference between the atoms of different elements? The difference

Proton: symbol = P or $H^+$
mass = 1 atomic mass unit (amu)
charge = +1

Neutron: symbol = N or n
mass = 1 amu
charge = 0

Electron: symbol = e
mass = 0
charge = -1

SLIDE #37




SLIDE #38

*(See Page 47A - Following)*

SLIDE #39

SCRIPT

SLIDES AND NOTES

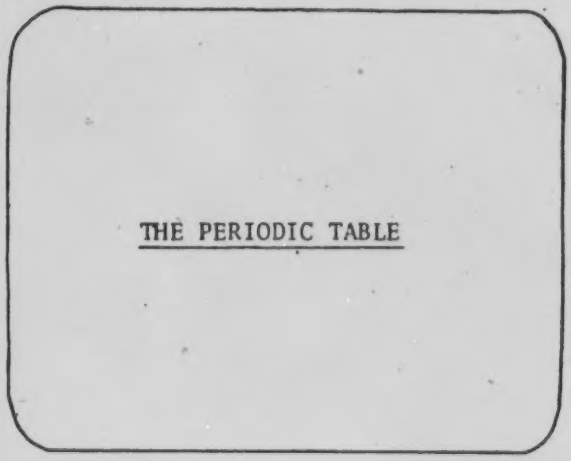
simply is that different elements have different numbers of protons, neutrons and electrons.

Hydrogen has but one proton and one electron. Helium, the second element of the periodic table has 2 protons, 2 neutrons, and 2 electrons. The third element, Lithium, has 3 protons, 3 electrons, and 4 neutrons, and so on.

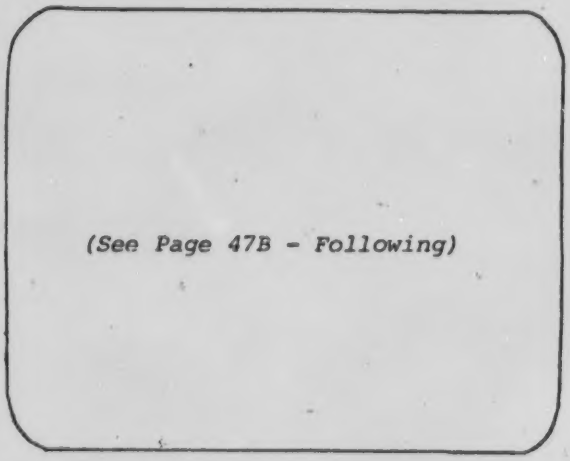
The atomic number of an element is the number of protons in the nucleus. The atomic mass is equal to the sum of the number of protons and neutrons. Electrons have negligible mass.

Advance your carousel to *Slide #40*. Notice that the atomic numbers are always whole numbers, while the weights are fractions. Does this mean that atoms can have fractions of neutrons in their nucleus? No, the reason the weights are fractions is that the number of neutrons in the atoms of elements may vary.

Look at *Slide #41*. Hydrogen may have 0, 1, or 2 neutrons in its



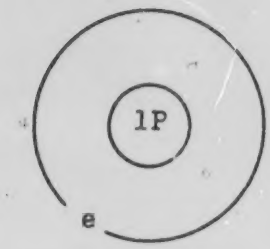

SLIDE #40



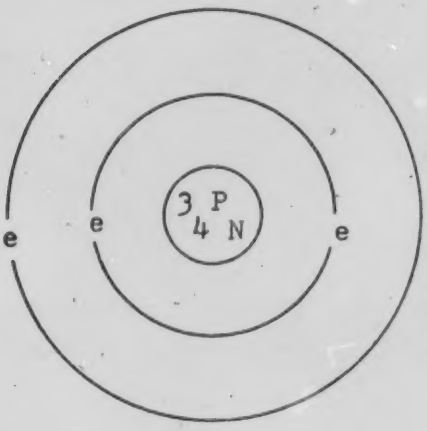

SLIDE #41

MODULE D1

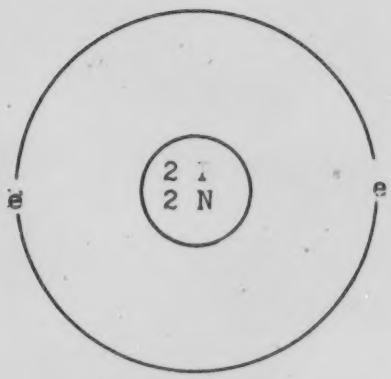
SLIDE #39




HYDROGEN




LITHIUM

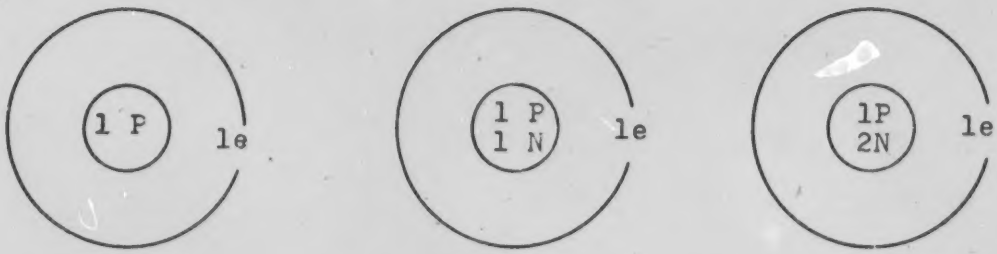



HELIUM

MODULE D1

SLIDE #41

ISOTOPES OF HYDROGEN




Hydrogen Deuterium Tritium

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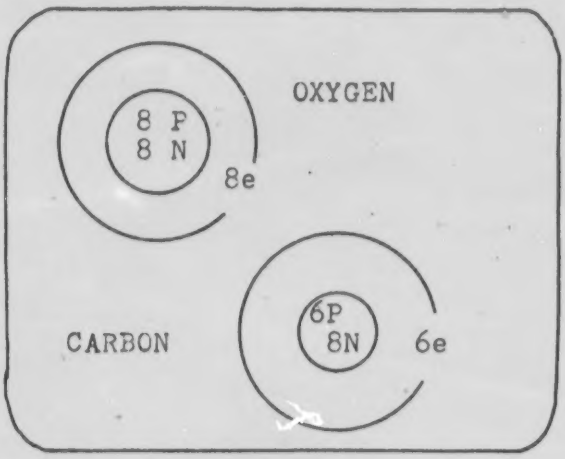
SLIDES AND NOTES

nucleus. The type of hydrogen with no neutrons is the commonest. The other two types are heavy hydrogen. The different forms of hydrogen are called isotopes. The 3 isotopes of hydrogen then are hydrogen, deuterium, and tritium.

Since all species of atoms have isotopes mixed together, their atomic weights are averages and therefore come out as fractions.

See if, on a piece of paper, you can draw a diagram of oxygen, atomic number 8 and atomic weight 16; and carbon, atomic number 6, atomic weight 14.

Look at *Slide #42*. You may have drawn these atoms as are shown in the slide. This, however, is not entirely correct. Electrons are distributed about the nucleus in shells. Oxygen (O), for example, has two shells. It has two electrons in its innermost shell or level and six in its outermost shell. Carbon (C) has two shells also. In its innermost shell it has




SLIDE #42

SCRIPT

SLIDES AND NOTES

two electrons while its outer shell contains four. If your original drawing was incorrect, draw it correctly now, and then, consult the section in the study guide titled "Orbitals."

You should note that the innermost ring or orbital has two electrons in the cases of both carbon (C) and oxygen (O). Orbitals have a limited capacity because electrons are all negative and therefore repel each other. This means that only so many can fit into a given space.

Look at *Slide #43*. The shells are lettered by convention. The innermost shell is called the K shell or orbital and the shells are alphabetically labeled from thereon. The capacities of the first three orbitals are given. Now consult number two in the orbital section of the study guide.

An atom of an element is electrically neutral, therefore, the number of protons and the number of electrons are equal.

Look at *Slide #44*. If an atom

*(See Page 49A - Following)*

SLIDE #43

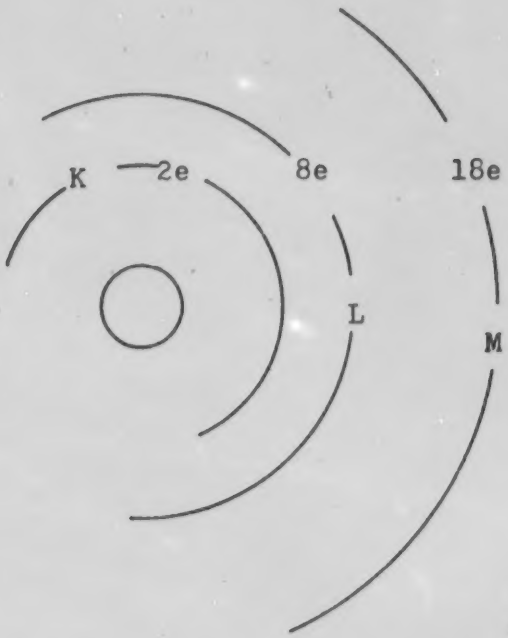
*(See Page 49B - Following)*

SLIDE #44

MODULE D1

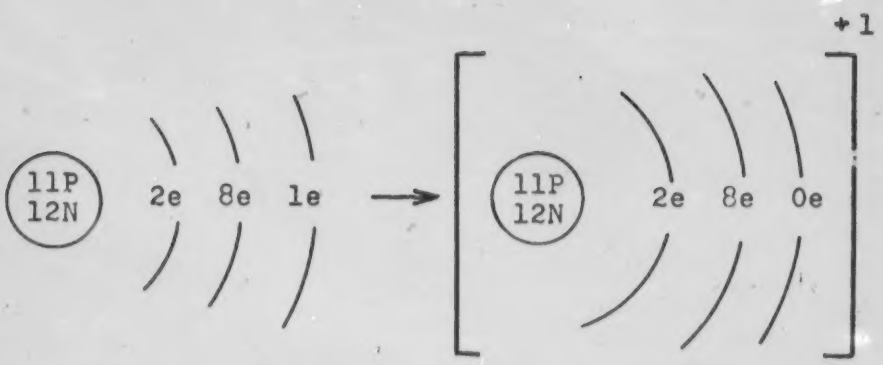
SLIDE #43

oxygen



MODULE D1

SLIDE #44




SODIUM ATOM MINUS ONE ELECTRON YEILDS A SODIUM ION

SCRIPT

gains or loses electrons it becomes charged and is then called an ion. The removal of 1 electron from a sodium atom leaves 11 protons and only 10 electrons, and therefore, a net charge of +1 on the resulting ion.

Now look at *Slide #45*. Chlorine has 17 protons, 18 neutrons, and 17 electrons. If it gains 1 electron it becomes a negatively charged ion.

Generally, a gain of electrons produces negative ions while loss of electrons produces positive ions. Metals tend to loose electrons, while non-metals tend to gain electrons. Hence, the difference between them.

In chemical reactions, electrons are lost, gained, or shared. There is never a change in the nucleus. All chemical reactions, from rock formations to the growth of your body, are the result of changes in the electronic structure of atoms.

Stop your tape and consult the "Atomic Structure" section of the mini-course guide.

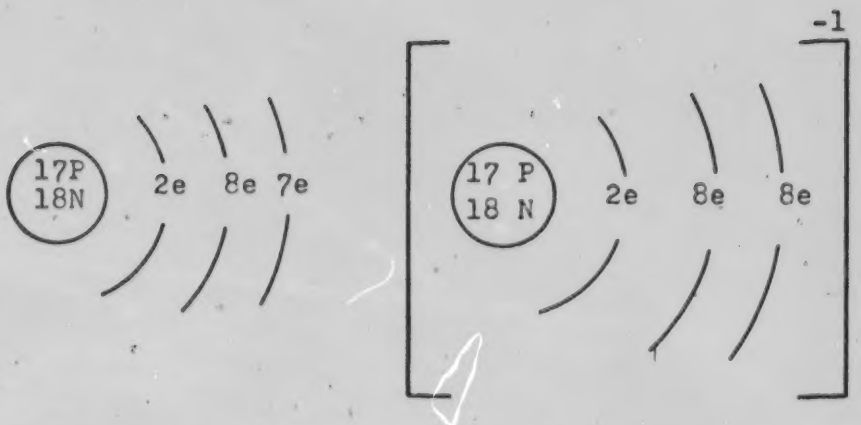
SLIDES AND NOTES

*(See Page 50A - Following)*

SLIDE #45

MODULE D1

SLIDE #45




Chlorine Atom

Chlorine Ion

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SLIDES AND NOTES

Your answers for questions numbers 2 and 3 should have been the following: for #2a. 36 atomic mass units or AMU; b. 98 AMU; c. 40 AMU; d. 68 AMU; e. 18 AMU; f. 32 AMU. For number 3a. 0; b. 10; c. 18; d. 10; e. 18. If you feel that you can draw the electron orbitals with some facility, you should continue on. In the event that you still have some questions, consult your student text. In addition, ask your instructor for the supplemental overhead projection sheets titled, "Atomic Structure." If you are enrolled in Ch-1C you will find these sheets in Appendix C of your mini-course study guide.

*(See Page 51A - Following)*

SLIDE #46

Advance to *Slide #46*. In this section we will consider bonding, how the atoms in compounds are held together. Some of the basic ideas are outlined on the slide. Chemical reactions do not involve all the electrons in the outermost shell are called valence electrons. Valence electrons are involved in chemical

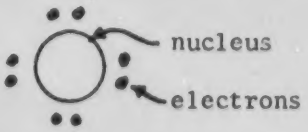
MODULE D1

SLIDE #46

The electrons in the outermost shells of an atom are called VALENCE electrons.

A valence shell is complete with 8 electrons.

Shells with 8 electrons are stable: those with less than 4 electrons tend to lose electrons, and result in metallic properties; those shells with more than 4 electrons tend to gain electrons, and result in properties we call non-metallic.

STABLE: 

SCRIPT | SLIDES AND NOTES

change.

The most stable electron configuration for a valence shell is 8 electrons. The exception is the K shell which is so small that it is full with 2 electrons. Only 2 elements, hydrogen and helium, have the K shell as their valence shell.

Atoms tend to gain or lose electrons to fill or eliminate their valence shells. Those atoms with only a few electrons tend to lose while those with nearly 8 electrons tend to gain.

For example, sodium has only one electron in its valence shell. To complete that shell, it would have to gain 7 more electrons, giving the ion a charge of -7 (a negative charge indicates that the ion possesses an excess of electrons while a positive charge indicates the ion has a deficiency of electrons). It is easier to rid itself of one electron and maintain a difference of +1 between the electron shells and the nucleus than

SCRIPT

it is to gain 7 electrons and maintain a -7 difference in charge. Conversely, it is easier for chlorine, which has 7 valence electrons, to gain one electron than to lose 7.

Atoms which tend to lose electrons have less than 4 valence electrons and are metals. The properties of metals, including luster, ductility, malleability, etc., and are due to the ease with which they lose valence electrons.

Atoms which tend to gain electrons are non-metals and have more than 4 valence electrons.

Elements which have exactly 8 valence electrons tend to neither gain nor lose electrons and are therefore inert.

Those atoms with 4 electrons usually do not gain or lose, but rather share electrons.

Stop your tape now and go to the "Valence" section of your study guide.

Look at *Slide #47*. Atoms which gain or lose electrons become ions. Ions are charged and, if their charges

SLIDES AND NOTES

*(See Page 53A - Following)*

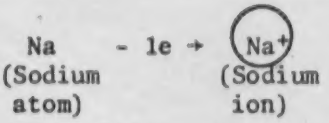
SLIDE #47

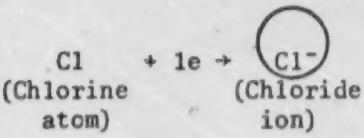
MODULE D1

SLIDE #47

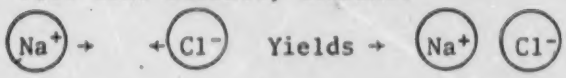
D1-47

Ionic bonding results when atoms, having gained or lost electrons, become charged and attract each other.




Free ions attract, combine.



SCRIPT

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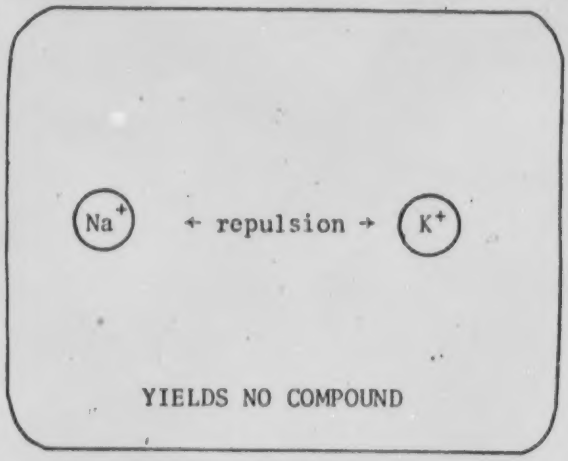
are opposite, they attract. Metals do not react with metals because they all form positive ions. Which of the atoms listed in the "Valence" section of the study guide will probably react or combine with one another and which species will not?

Advance your carousel to *Slide #48*. The ionic bonds that result from a combination of ions do not produce molecules, but instead crystals with no discernable molecules.

Advance your carousel to *Slide #49*. Slide #49 depicts a salt crystal. The structure is such that it cannot be considered a molecule unless one were to consider the whole crystal as a molecule. There is nothing in this crystal that could be considered a molecule unless it is the whole crystal.

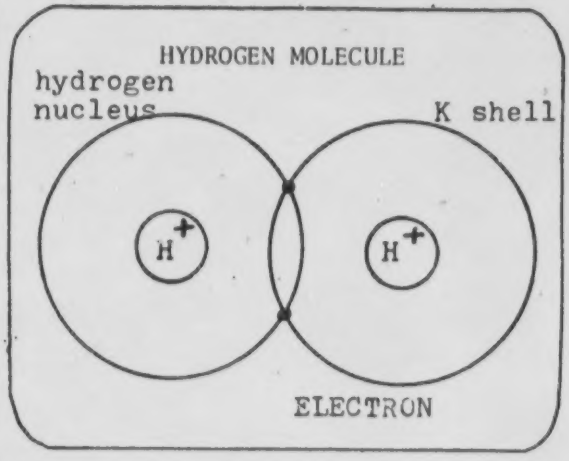
Look at *Slide #50*. Atoms can sometimes more easily fill their shells by sharing electrons. A bond formed by sharing electrons is called a covalent bond. Notice that a dif-




SLIDE #48

Slide shows: a picture of a cubic of NaCl showing the relative locations of $Na^+$ and $Cl^-$.

SLIDE #49




SLIDE #50

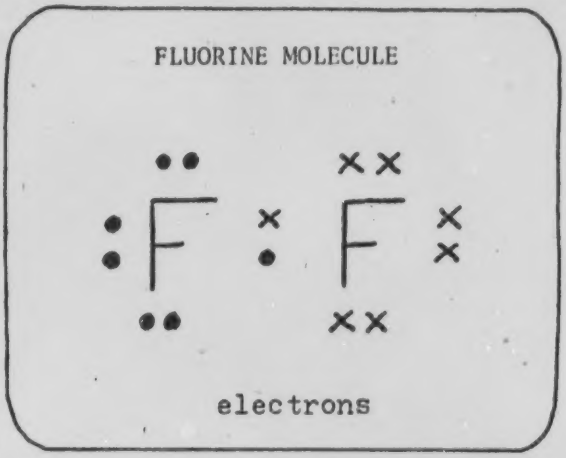
SCRIPT

SLIDES AND NOTES

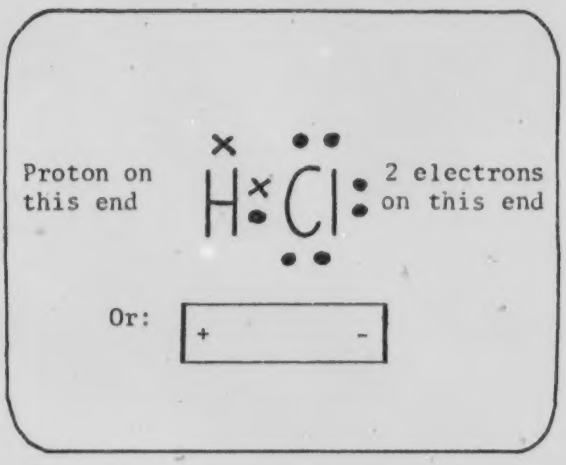
ference in charges still holds the atoms together. The two positive hydrogen nuclei are drawn to the relatively negative area between them caused by the shared electrons. Sharing electrons also completes the valence shell of each atom.

Look at *Slide #51*. Only valence electrons are shown, we ignore the others. The letter F represents the nucleus. The symbols of the electrons are as shown. Note that by sharing electrons, each atom can be surrounded by 8 valence electrons.

Look at *Slide #52*. Covalently bonded molecules are quite often polar, that is, they have a small charge on each end. One end of the hydrogen chloride (HCl) molecule has a hydrogen (H) or proton protruding. Protons, you will recall, are positive. On the other end electrons are protruding, therefore, that end is negative. The molecule has a positive and a negative pole and, thus, is called a polar molecule.




SLIDE #51




SLIDE #52

SCRIPT

Look at *Slide #53*. Water molecules are not straight, but are bent. This makes them polar. Being polar, they tend to stick together and are hard to pull apart. This is the reason why it takes so much energy to evaporate water.

Look at *Slide #54*. Electrovalent bonds form when atoms gain or loose electrons. Covalent bonds on the other hand form when atoms share electrons. The number of electrons shared completes the outer shell of each atom. Molecules are discrete neutral particles formed by covalently bonded atoms; ionic bonds, on the other hand do not form molecules. Consult the bonding section of the study guide.

Advance your carousel to *Slide #55*. Consider what happens when a material dissolves. The simplest thing that can happen is that the molecules of solvent and solute just mix. For example, when alcohol and water mix, or when oxygen dissolves in water, that is about what happens.

SLIDES AND NOTES

Slide shows: a picture showing the hydrogen bonding between several water molecules.

SLIDE #53

Electrovalent bond = ionic bond.

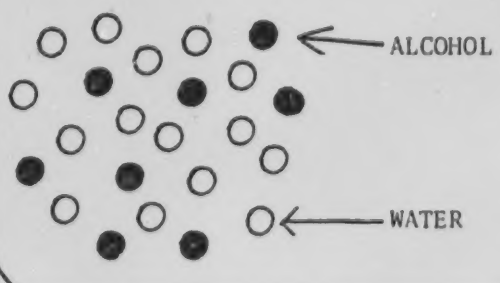
Particles which have gained or lost electrons (ions) form electrovalent bonds and result in ionic crystals.

Covalent bonds result from sharing electrons.

Covalent bonds result in molecules.

SLIDE #54

MODEL OF AN ALCOHOL-WATER SOLUTION:




SLIDE #55

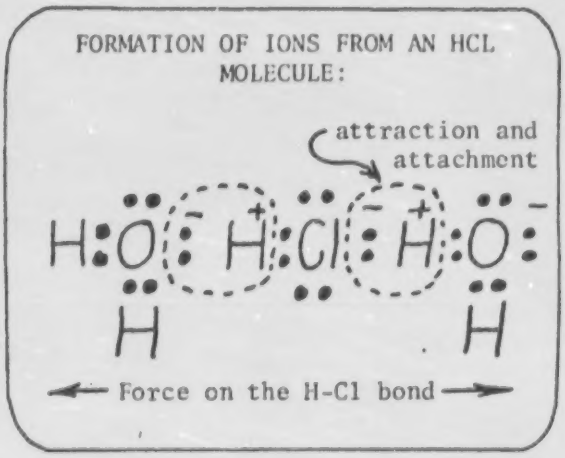
SCRIPT

SLIDES AND NOTES

For most solutions, the picture is somewhat more complex. The best solvents, such as water ($H_2O$), are polar covalent materials.

Look at *Slide #56*. The polar nature of water helps it tear apart materials so that soluable ions form. Polar water molecules can pull apart a polar covalen. molecule, such as hydrogen chloride (HCl) to yield a hydrogen ion (H+) and a chlorine ion (Cl-). The process is called ionization. Ionization then is the production of ions from molecules.

Look at *Slide #57*. In the case of ionic substances, a similar mechanism dissolves the solute. The surface ions, and particularly the corner ions of a crystal are subject to attachment by a water molecule. As the molecules in the liquid vibrate and move, and the ions in the crystal vibrate, the water molecule often succeeds in pulling off an ion. As a result, the crystal dissolves. This process is called dissociation. Dissociation is




SLIDE #56

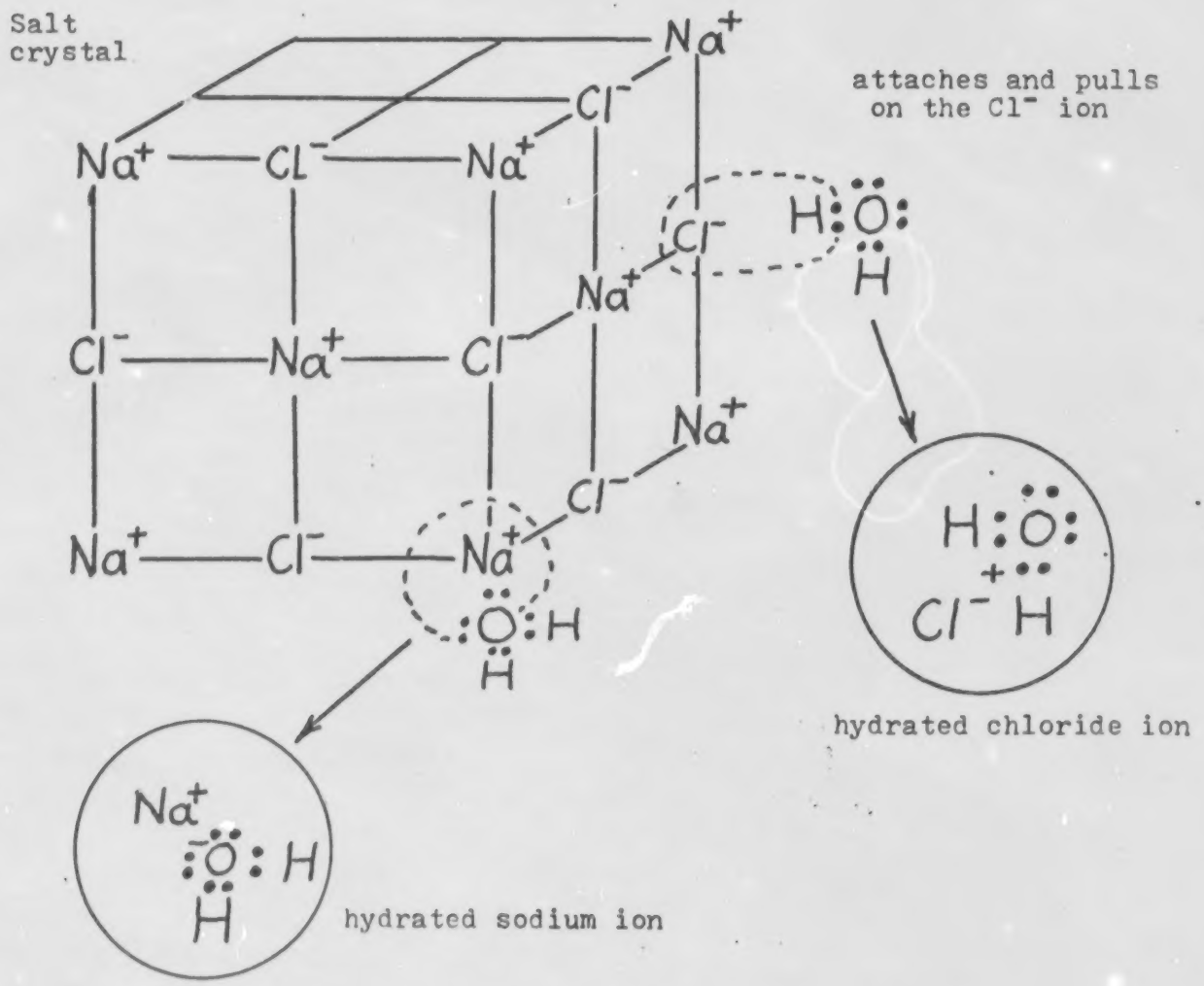
*(See Page 57A - Following)*

SLIDE #57

MODULE D1

SLIDE #57

D1-57 DISSOCIATION OF AN IONIC MATERIAL



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SLIDES AND NOTES

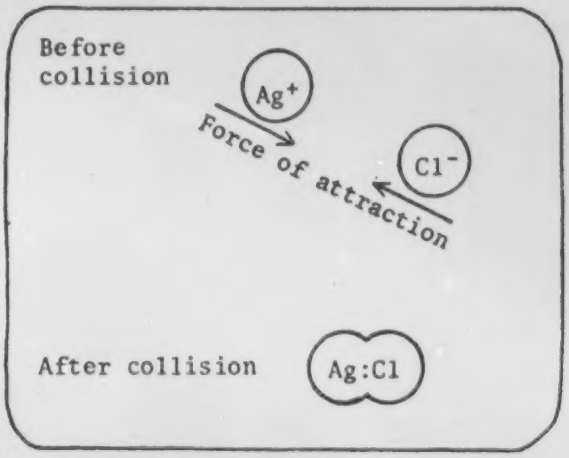
the separation of ions from an ionic solid.

At this point, one questions how crystals form in solution.

One might ask how do crystals form in solution. Crystals stop dissolving when sufficient numbers of ions are floating around in solution so that the chance of an ion being pulled off the crystal equals the chance of it being replaced. This point is called the saturation point.

Crystals form when the probability of an ion being replaced is greater then the probability of an ion being pulled off the crystal. This is why evaporation of a liquid, which increases the concentration of ions, causes dissolved salts to precipitate.

Look at *Slide #58*. Chemical reactions occur in solution if the result is removal of ions from solution. Ions in a solution move about at random and experience thousands of collisions per second. Sometimes, a pair of ions collide and form a bond strong




SLIDE #58

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enough that the solvent doesn't pull them apart. On meeting a chloride ion (Cl), a silver ion (Ag) forms a stable molecule of silver chloride (AgCl). These molecules collect and form a visable white solid which is called a precipitate.

This principle is used in boiler water analysis, in ascertaining the salinity of sea water and for many other purposes. Under controlled conditions, as was the case with acid-base neutralization, it is called titration. As before, the important point is the end point or that point where all of the chloride ion has been removed from solution. Turn now to Appendix A of your minicourse guide and conduct lab exercise #7.

Advance to *Slide #59*. This particular reaction is used to analyze for chloride ions in boiler water.

One of the concerns of operating marine engineers is the formation of sludges. A sludge, as you perhaps now realize, is a precipate.

Slide shows: a bottle containing sea water; one containing $AgNO_3$; and a 125 ml. Erlenmeyer flask.

SLIDE #59

SCRIPT

As you reflect upon your study to this point, can you explain ionization, dissociation, how crystals are dissolved, and how precipates are formed? Try these on a separate sheet of paper and then compare your answers with those provided on the audio tape.

Advance your carousel to *Slide #60*. A chemical symbol can be used to represent the name of an element, but in a formula the chemical symbol represents 1 atom of that element. A formula, made up of symbols, represents the composition of one molecule, or in the case of an ionic substance, the symbol represents the simplest ratio of atoms. Subscripts indicate the number of atoms in the molecule.

Look at *Slide #61*. Number 1 is the formula for water ($H_2O$). A molecule of water consists of 2 atoms of hydrogen (H) and one atom of oxygen (O). Number 2 is iron oxide ($Fe_2O_3$). A molecule of iron oxide contains 2 atoms of iron (Fe) and 3 atoms of oxygen (O).

SLIDES AND NOTES

A chemical symbol represents one atom of an element.

A formula represents the composition of one molecule of a substance - or the simplest ratio of atoms in an ionic substance.

SLIDE #60

1. $H_2O$
2. $Fe_2O_3$
3. $MgCO_3$
4. $H_2SO_4$

SLIDE #61

SCRIPT

What are the contents of the molecules 3 and 4? Stop the tape a minute and figure it out.

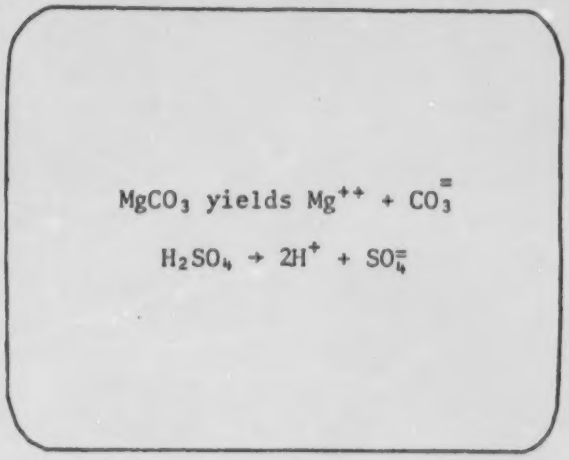
Number 3 is magnesium carbonate ($MgCO_3$), a common scale formed in boilers. It contains one magnesium atom (Mg), 1 carbon (C), and 3 oxygen atoms (O).

Number 4 is sulfuric acid ($H_2SO_4$). It contains 2 hydrogens (H), 1 sulfur (S), and 4 oxygen atoms (O).

Advance your carousel to *Slide #62*. In molecules that have more than 2 kinds of atoms, the bond strength between different species of atoms is usually unequal. Slide #62 shows what happens to the two molecules with which we were just working, when they break up.

Magnesium carbonate ($MgCO_3$) breaks between the magnesium (Mg) and the carbon (C) rather than between carbon (C) and oxygen (O). Therefore, the two fragments that result on breakup are a magnesium ion (Mg) and a carbonate ion ($CO_3$).

SLIDES AND NOTES




SLIDE #62

SCRIPT

Groups of atoms that are bonded strongly enough to go through chemical reactions together are called radicals.

When something breaks up a sulfuric acid molecule ($H_2SO_4$), the hydrogens (H) become separated. The 3 resulting fragments are 2 hydrogen ions and a sulfate radical ($SO_4$). The sulfate radical is charged and so is also an ion.

Look at *Slide #63*. Some other radicals frequently associated with boiler water treatment are listed.

Advance to *Slide #64*. A chemical equation consists of formulas which indicate the reactants and products of a reaction.

Look at *Slide #65*. It is convention to write the formulas for the materials that react on the left side and the results of the reaction, or products, on the right side. The coefficients in front of formulas indicate the number of molecules involved in the reaction.

The first equation states that 2

SLIDES AND NOTES

SOME OTHER RADICALS:

| | | | |
|---|---|---|---|
| $CrO_4^{-2}$ | chromate | $SiO_3^{-2}$ | silicate |
| $NO_3^{-1}$ | nitrate | $PO_4^{-3}$ | phosphate |
| $NO_2^{-1}$ | nitrate | $HCO_3^{-1}$ | bicarbonate |
| $SO_3^{-2}$ | sulfite | $NH_4^{+1}$ | ammonium |
| $OH^{-1}$ | hydroxide | | |

SLIDE #63

A chemical equation gives an overall description of a chemical reaction. It consists of formulas which indicate the materials going into the reaction (REACTANTS) and the materials coming out of the reaction (PRODUCTS).

REACTANTS —yield→ PRODUCTS

$A + B + C + \ldots \rightarrow X + Y + \ldots$

$2H_2 + O_2 \rightarrow 2H_2O$

SLIDE #64

1. $2H_2 + O_2 \rightarrow 2H_2O$
2. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
3. $2Ag^+ + CaCl_2 \rightarrow 2AgCl + Ca^{++}$

$Ag^+$ = SILVER ION

$CaCl_2$ = CALCIUM CHLORIDE

$Ca^{++}$ = CALCIUM ION

AgCl = SILVER CHLORIDE

SLIDE #65

| SCRIPT | SLIDES AND NOTES |
| --- | --- |
| molecules of hydrogen (H) gas, each containing 2 atoms of hydrogen (H) react with one molecule of oxygen (O), containing 2 atoms, to yield 2 molecules of water ($2H_2O$), each containing 2 atoms of hydrogen (H) and one of oxygen (O).<br><br>The second equation states that one atom of zinc (Zn) containing 1 atom per molecule, reacts with 2 molecules of hydrogen chloride (HCl) to yield one molecule of zinc chloride (ZnCl) plus one molecule of hydrogen (H).<br><br>Reaction number three states that two silver (Ag) ions react with 1 calcium chloride (CaCl) molecule to yield 2 silver chloride (AgCl) molecules and a calcium (Ca) ion.<br><br>Now consult the "Chemical Equation" section of the minicourse study guide. | |

# APPENDIX E

# EXAMINATIONS

MINICOURSE D-1-1 EXAMINATION

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. A container of liquid is found. Its density is listed on the container as 1 Kg/liter. The mass of the container plus the liquid is 50 Kg and the mass of the container is 25 Kg. How much of the liquid has been found.
   a. 50 l
   b. 25 ml
   c. 2,500 ml
   d. 500 ml
2. A sample of water is drawn from a ship's boiler and phenolphthalein added. This apparently has no effect since the sample remains clear. From this observation the tester should assume that the boiler water:
   a. is slightly basic
   b. has a ph of 13
   c. contains chlorine
   d. is acidic
3. Elements at the extreme right side of the periodic chart have 8 electrons in their outer shells. This makes them:
   a. gases
   b. inert
   c. transition elements
   d. solids

4. A covalent bond can be defined as:

   a. a bond in which atoms share electrons.

   b. a bond in which electrons are transferred from one atom to another.

   c. the result of mixing together more than three reactants.

   d. a bond showing the same properties as an ionic bond.

Match the following: a letter may be used more than once.

| | |
|---|---|
| 5. ion | A. $H^+$ |
| 6. electron | B. The smallest subdivision of matter which identification of the matter. |
| 7. element | C. A negatively charge particle whose mass is minute. |
| 8. neutron | D. 92 naturally occurring ones are known to man. |
| 9. proton | E. A neutrally charged nuclear particle whose mass is IAMU. |

10. A solution of water and some other species is placed on a warm stove. What would you expect to happen to the solvent/solute ration.

    a. increase

    b. decrease

    c. remain the same

    d. explode

Identify the following as: (A) formulae; (B) radicals; (C) ions; (D) equations.

11. $SO_4^=$

12. $H_2O$

13. $Na^+$

14. $C_6H_{12}O_6$

15. $Hcl + NaOH \rightarrow H_2O + NaCl$

16. Describe how a burette is used.

17. Describe a step-by-step procedure which would lead you to state that all of the chlorine had been removed from a water sample.

MINICOURSE D-1-2 EXAMINATION

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. Define each of the following:
   a. Density
   b. Solution
   c. Solvent
   d. Element
   e. Acid
   f. Neutralization
   g. Organic Compound
   h. Atom
   i. Isotope
   j. Ion
   k. Ionic Bond
2. Draw a sketch of the Periodic Table giving the location of metals, nonmetals, and inert gases. Also, explain the reasons why the elements are placed where they are. Use the back of the sheet for your discussion.

3. Draw a model of the carbon atom. Carbon = At. Wt. 12 and At. #6. Show the locations of the electrons.

4. Ascertain the mass of a sample whose density is 11.34 g/cc and whose volume is 77 cc.

5. Give an example of a formula and a radical.

6. How many valence electrons do each of the first 10 elements of the periodic table have?

7. You observe a small leak in a steam line. A drop of water constantly drops to a metal block below, and the spot where it hits seems to be corroding. If you took a sample of steam line water and added phenolpthalein what would you expect to observe?

MINICOURSE D-1-3 EXAMINATION

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. Give two examples of elements.
2. Give two examples of mixtures.
3. Give two examples of compounds.
4. Give two examples of suspension.
5. Make a rough sketch of the Periodic Table and show the location of the METALS, NON-METALS, METALOIDS, INERT GASES, and discuss in detail the reason why elements are placed where they are. Use the back of this sheet.
6. Flourine (F) has an atomic number of 9 and an atomic weight of 19.
   a. How many protons does it have? _____.
   b. How many neutrons does it have? _____.
   c. How many electrons are there in the neutral atom? _____.
7. There is only one thing which really distinguishes an atom of one element from another atom of another element. It is the number of _____.
8. An ion of an element is an atom of that element which has _____.
9. An isotope of an element is an atom of that element which has _____ _____.
10. Draw a diagram of an atom of Sodium (Na), which has an atomic number of 11 and an atomic mass of 22.99, and show the electrons in each of the shells.

11. Define the following terms:

    a. formula

    b. equation

    c. radical

12. Compute the mass (in grams) of a rectangular solid whose dimensions are 10 cm x 100 mm x 1 clm, and whose density is 1 kg/1,000 cc.

13. How many valence electrons have each of the first ten elements of the periodic table.

14. Would you expect these elements (from Question #13) to gain or lose electrons if they were to become ionized?

15. Discuss the ways you have thus far used chemical indicators. What have they told you?

MINICOURSE D-1-4 EXAMINATION

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. An atom of oxygen has the atomic weight of 17 and an atomic number of 8.
   a. How many electrons does it have? _____.
   b. How many protons does it have? _____.
   c. How many neutrons does it have? _____.
   d. How many electrons does it need to fill up its outer shell? _____
   _____.
2. Sketch the Periodic Table and show the location of the metals, nonmetals, metaloids, and noble gases.

3. Is brass a mixture or a compound? Explain your answer. _____
   _____
   _____
   _____
4. What is the difference between a true solution (e. g., salt water) and a suspension (e. g., muddy water)? _____
   _____
   _____
   _____

5. When an acid and a base neutralize, what are the products? _____ and _____.

6. What is it about inert gases (noble gases) that makes them inert (unable to react with anything else)? _____

_____.

7. Name the two types of chemical bonding and explain the difference.

   a. _____ bond, electrons are _____.

   b. _____ bond, electrons are _____.

8. Define the following terms:

   a. solvent _____.

   b. titration _____.

   c. solute _____.

   d. indicator _____.

   e. density _____.

9. Calculate the density of a liquid whose mass is 35 g and is held in a completely filled container with a height of 10 cm. and a diameter of 20 cm.

10. Diagram the valence electrons in the following molecules: $H_2O$, Hcl, $CH_4$, $NH_3$.

11. What differences would you expect to find if you neutralized .5M NaOH with .5M $H_2SO_4$ rather than .5M Hcl?

12. Describe the correct method of reading the fluid level in a graduated cylinder. _____

_____

_____

_____.

13. When titrating for salinity, what is the meaning of salinity? _____

_____

_____

_____.

MINICOURSE D-1-5 EXAMINATION

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. An object has a volume of 35 cubic centimeters and a mass of 17.5 grams. What is its density? _____.

2. Diagram an atom having three protons, two neutrons, and three electrons.

3. Define the following:

   a. Solvent _____.

   b. Solute _____.

   c. Solubility _____.

   d. Salt _____.

   e. Phase of matter _____

   _____.

   f. Ion _____.

   g. Electron _____

   _____.

4. What are the products of an acid-base neutralization? _____

   _____

   _____.

5. Describe the construction of the Periodic Table. In your discussion include the reasons why an element is placed in one spot and not another. _____

_____

_____

_____

_____

_____

_____

_____

_____

_____.

6. Identify each of the following as either formula, radical, ion, or atom:

a. $SO_4$ = _____ d. $H^+$ _____

b. $H_2SO_4$ _____ e. S _____

c. $K_2CrO_4$ _____ f. $HCl + NaOH \rightarrow NaCl + H_2O$ _____

7. List the number of valence electrons in elements whose atomic numbers are 10 - 16.

8. Describe the use of a burette in an acid-base neutralization. _____

_____

_____

_____

_____.

MINICOURSE D-1-6 EXAMINATION

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. Aluminum has a density of 2.70 g/ml. What is the mass of a rectangular block 10 cm. x 100 mm x $10^{-1}$ M?

2. Describe what happens as the energy of molecules increases. _____

_____

_____

_____

_____.

3. Sketch a picture of the Periodic Table. Describe its construction and the factors determining the placement of the elements.

4. Define each of the following:

   a. Radical _____

   _____

b. Equation _____

_____.

c. Formula _____

_____.

d. Precipitation _____

_____.

e. Dissociation _____

_____

f. Covalent Bond _____

_____.

g. Neutron _____

_____.

h. Ion _____

_____.

5. You are shown a section of pipe through which water constantly passes. During your inspection you note that the inner surface seems to be eroding away. Describe the laboratory procedures that will allow you to evaluate this situation. _____

_____

_____

_____

_____

_____.

6. What are end points? How do you know when you have found one? _____

_____

_____

_____

_____.

7. Give an example of an equation, a radical, and an ion.

8. Make a chart showing the electrons in the orbitals of the first ten elements.

MINICOURSE D-1-7 EXAMINATION

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. Define each of the following:

   a. Colloid _____

   _____

   b. Neutralization _____

   _____

   c. Charge _____

   _____

   d. Ionic Bond _____

   _____

   e. Isotope _____

   _____

   f. Proton _____

   _____

2. Draw a picture of the Periodic Table, showing the placement of inert gases, metals, etc.

3. Explain the reasons for the placement of elements in their respective places on the Periodic Table _____

4. If atoms gain or lose electrons in the bonding process, what type of bond is formed? _____

5. How many atoms of each element are there in the following formula $K_2CrO_4$?

6. A liquid has a density of 2.2g/cc and a mass of 300 g. If the liquid is to be stored in a container whose diameter is 8 cm., how tall must the container be?

7. Draw pictures of the valence electron configurations for the following molecules (K = 19, Cl = 17, C = 6, O = 8, H = 1, N = 7, Na = 11):

   a. KCl

   b. CO

   c. $CO_2$

   d. $H_2O$

   e. $NH_3$

   f. NcCl

8. Describe in detail how you might determine that all of the chlorine had been removed from a sample of water. _____

9. What is meant by acid-base neutralization, and how might the process be used to advantage by a merchant mariner at sea?

MINICOURSE D-1-8 EXAMINATIONS

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. When in chemistry are salt and water formed? _____

_____

_____

_____

2. Describe the Law of Definite Proportions and how it relates to chemistry.

_____

_____

_____

3. Draw a diagram of an atom of magnesium, which has an atomic number of 12 and an atomic mass of 24.3. Show the electrons in each of the shells.

4. Define each of the following:

   a. Organic compounds _____

   _____

   b. Homogeneous _____

   _____

   c. Indicator _____

d. Electrovalent bond _____

e. Dissociation _____

5. Sketch the Periodic Table, showing the placement of inert gases, metals, etc. Discuss the reasons why elements are placed where they are on the table.

6. What is a burette and how is it used? _____

7. A sample of liquid is poured into a graduated cylinder. Explain how you determine the exact amount of liquid in the cylinder. _____

8. What is the meaning of valence electron and how many have elements numbers 6 through 12. _____

9. Give an example of each of the following: atom, ion, formula, radical, equation, element.

10. Why can there be but two electrons in the first shell? _____

MINICOURSE D-1-9 EXAMINATION

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. How is an atom of an element represented? _____

_____

2. How is the mass of an atom determined? _____

_____

3. The Periodic Table contains a group of elements called the inert gases. Discuss the characteristics peculiar to these gases. _____

_____

_____

_____

_____

_____

4. Sketch the Periodic Table showing the placement of metals, nonmetals, inert gases, etc. Discuss the reasons why these groups are placed where they are.

_____

_____

5. An atom of oxygen has an atomic mass of 15.99 and an atomic number of 8. Make a sketch of that atom showing numbers of protons, neutrons, and electrons and their placement.

6. Define each of the following terms:

   a. Solute _____

   b. Solid _____

   c. Density _____

   d. Evaporation _____

   e. Homogeneous _____

   f. Colloid _____

7. 500 ml. of chlorine gas at $0^0$ C and 1 atmosphere has a mass of 50 mg. What is the density of the gas in grams per liter at conditions of standard temperature and pressure?

8. How many AMU's are there in each of the following (H = 1, Cl = 35, S = 32, O = 16, K = 39, Cr = 52, Na = 23):

   a. HCl _____ c. NaCl _____

   b. $H_2SO_4$ _____ d. $H_2O$ _____

   c. $K_2CrO_4$ _____ e. NaOH _____

9. Give an example of each of the following: radical, equation, formula, ion, element, and an atom.

10. A ship's drinking water supply is suspected of being contaminated with seawater. Discuss in detail the steps you might take to either confirm or refute your suspicions. _____

11. Discuss the steps you would take in an acid-base titration, the lab equipment you would use, and the role of the indicator. _____

MINICOURSE D-1-10 EXAMINATION

NAME _____

STUDENT ID NUMBER _____

DATE _____

1. Define each of the following:

   a. Density _____

   _____

   b. Acid _____

   _____

   c. Malleability _____

   _____

   d. Neutron _____

   _____

   e. Ionic bond _____

   _____

   f. Formula _____

   _____

   g. Mixture _____

   _____

   h. Periodic table _____

   _____

   i. Organic chemistry _____

   _____

   j. Elements _____

   _____

k. Electrons _____

_____

l. Atoms _____

_____

m. Neutralization _____

_____

2. Sketch the Periodic Table, showing the placement of metals, nonmetals, inert gases, etc. Also, describe what determines the placement of the various elements.

_____

_____

_____

_____

_____

_____

3. Draw a sketch of an atom of flourine (F), atom mass = 18.99840 and atomic number = 9.

4. The mass of a container is 170g and the mass of the container plus its liquid is 193.7g. If the volume of the fluid is 30 ml. and the density of seawater is 1.025g/cc, is it possible that the liquid in the container is seawater? _____

5. Make sketches of elements numbers 8 through 14. Include in your sketches the protons, neutrons, and location of electrons.

6. Liquid A is observed to be floating on the surface of Liquid B. What statement can you make concerning the densities of these two liquids? _____

_____

_____

_____

_____

7. What is meant by the law of definite composition; how might you demonstrate the phenomenon to a group of students; and what is its relationship to $H_2SO_4$? _____

_____

_____

_____

_____

_____

_____

_____

_____

_____

8. What are titrations, and how might they be used by the merchant mariner?

9. Name 5 pieces of glassware that you might find in a general chemistry laboratory and give one use for each piece.

10. Discuss a use of silver nitrate in general chemistry which might also prove of use to merchant sailors.

11. Give the number of valence electrons for each of the first 16 elements.

12. Give examples of each of the following: indicator, negative ion, formula, radical, equation, and an atom.